

Atomic Structure

The section of the general chemistry course that deals with atomic structure is designed to describe in qualitative terms the current understanding of the fundamental structure of all matter. Here, atoms are introduced as the basic building blocks of matter. The composition of the atom is described. Frequently, the fundamental experiments that led to the discovery of electrons, protons, and neutrons are laid out in some detail.

The development of quantum mechanics enabled chemists to describe electron energies and locations outside the nucleus more accurately than was possible with the planetary model for the atom. The meanings and implications of quantum numbers, photons, electromagnetic radiation, and radial probability distributions are central to describing the atom in terms of quantum mechanics. Other central ideas include the *aufbau* principle and the uncertainty principle.

Symbolic representations of atomic structure and three-dimensional representations of electron probability regions enable chemists to deal with the behavior of atoms in terms of current atomic theory and to communicate chemical information more clearly.

Study Questions

AS-1. An atom of strontium-90, $^{90}_{38}\text{Sr}$, contains

- (A) 38 electrons, 38 protons, 52 neutrons.
- (B) 38 electrons, 38 protons, 90 neutrons.
- (C) 52 electrons, 52 protons, 38 neutrons.
- (D) 52 electrons, 38 protons, 38 neutrons.

Knowledge Required: (1) The symbolic representation of isotopes, with the atomic number of an element written as a left-hand subscript, the mass number written as a left-hand superscript, and the ionic charge (if any) written as a right-hand superscript. (2) The fact that the mass number is the sum of the number of protons and neutrons.

Thinking it Through: The atomic number of 38 indicates 38 protons, eliminating choice (C). This is a neutral atom, so there are also 38 electrons, eliminating choice (D). The mass number of 90 for strontium eliminates choice (B), which corresponds to a mass number of 128. The correct number of neutrons is found by subtracting the atomic number (38) from the mass number (90) giving a value of 52, which is found in choice (A). *Note:* There is redundant information in this question; the term "strontium-90" gives the same information as the symbolic representation, $^{90}_{38}\text{Sr}$ if you know the symbol for strontium and can find its atomic number on the periodic table.

AS-2. What do these have in common?

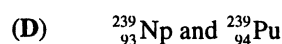
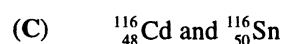
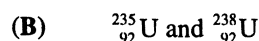
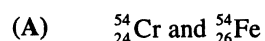


- | | |
|----------------------------------|---------------------------------|
| (A) the same number of protons | (B) the same number of neutrons |
| (C) the same number of electrons | (D) the same size |

Knowledge Required: (1) The symbolic representation of isotopes, with the atomic number of an element written as a left-hand subscript, the mass number written as a left-hand superscript, and the ionic charge (if any) written as a right-hand superscript. (2) The fact that the mass number is the sum of the number of protons plus neutrons.

Thinking it Through: The periodic table contains the atomic number of each element, which indicates both the number of protons in the nucleus and the number of electrons for a neutral atom. Negative ions have one or more additional electrons, and positive ions are short by one or more electrons. Also, subtracting the atomic number from the mass number gives the number of neutrons. These three species have different atomic numbers (10, 9, and 12), so the number of protons cannot be the same, eliminating choice (A). Subtracting each atomic number from the given mass number (20, 19, and 24) gives different numbers of neutrons in each case (10, 10, and 12), eliminating choice (B). Considering the number of electrons, the neutral atom neon has 10; fluorine has 9 in a neutral atom—and 10 in a fluoride ion; magnesium has 12 in a neutral atom—and 10 in a +2 magnesium ion. The number of electrons is thus the same for all three. Such species are called *isoelectronic*. Note that Na^+ , O^{2-} , and Al^{3+} are also isoelectronic with the given species. Choice (D) cannot be true; because, given equal numbers of electrons, those nuclei with greater positive charge will attract electrons more strongly than those with less positive charge.

AS-3. Which pair represents isotopes?



Knowledge Required: The definition of “isotope,” and the associated symbolic representation.

Thinking it Through: Isotopes are forms of the same element that differ only in the number of neutrons in their nuclei. The identity of an element is determined by the number of protons (the atomic number), which is represented by the left-hand subscript. Choice (B), the only choice with the same element symbol for both atoms, also has the identical atomic number (92). The other three choices have identical mass numbers, which simply means that the sum of protons plus neutrons is the same for each one. *Note:* The similar terms, *isotope*, *isomer*, and *allotrope* are often confused.

AS-4. Which particle, if lost from the nucleus, will *not* result in a change in the atomic number?

(A) proton

(B) alpha particle

(C) beta particle

(D) neutron

Knowledge Required: (1) The general composition of atomic nuclei (numbers of neutrons and protons). (2) The modes of radioactive decay available to nuclei. (3) The meanings of the terms alpha particle and beta particle.

Thinking it Through: Any of these particles could be lost from the nucleus during a nuclear (not chemical) change. Since the atomic number is determined by the number of positively charged protons in the nucleus, emission of *any* charged particle would change the atomic number. A proton, choice (A), has one unit of positive charge. An alpha particle, choice (B), has two units of positive charge. A beta particle, choice (C), has one unit of negative charge. The only remaining choice is the neutron, which has no charge, and therefore does not change the atomic number.

AS-5. When alpha particles were shot at a metal foil target to probe the structure of the atom, most of the particles passed through without any deflection in path. Some particles were deflected at large angles. This indicated to Rutherford that

- (A) the metal foil was continuous matter.
- (B) the mass of the metal atoms was spread out evenly.
- (C) the atoms of the metal were mostly empty space.
- (D) the alpha particles had great penetrating power.

Knowledge Required: The general features of Ernest Rutherford's alpha-particle-scattering experiment.

Thinking it Through: Prior to Rutherford's classic scattering experiment, atoms were thought to consist of a cloud of positive charge with electrons embedded throughout. Rutherford discovered that some alpha particles were scattered at large angles. The only explanation was that most of the mass and all of the positive charge had to be concentrated in a tiny fraction of the volume of atom. The rest of the atom had to be mostly empty space. Choice (B) describes atomic theory *before* Rutherford's experiment. Choice (A) suggests that atoms do not exist at all. Choice (D) does not bear directly on Rutherford's experiment. The large scattering angles of the alpha particles were *only* consistent with an atomic model in which the positive charge and almost all of the mass were concentrated in a small, dense center. The rest of the atom had to be mostly empty space, choice (C).

AS-6. An atom has a valence shell electron configuration of ns^1 . To which group of elements in the periodic table does it belong?

- | | |
|---------------------------|-------------------|
| (A) transition elements | (B) lanthanides |
| (C) alkaline earth metals | (D) alkali metals |

Knowledge Required: (1) The symbolic representation of electron configuration. (2) The meaning of the term "valence shell." (3) The names associated with particular families or periods in the periodic table. (4) The ability to correlate symbolic representations with locations in the periodic table.

Thinking it Through: Elements having ns^1 as the valence-shell-electron configuration have a single electron in the s orbital. Such elements include H ($1s^1$), Li ($2s^1$), Na ($3s^1$), and K ($4s^1$). These are the **alkali metals**, found in Group IA of the periodic table, choice (D). The alkaline earth metals, Group IIA, all have electron configurations of ns^2 , choice (C). Transition elements, choice (A), and lanthanides, choice (B), have varying numbers of valence electrons. Their characteristic valence electron configurations are $ns^2(n-1)d^x$ and $ns^2(n-1)d^x(n-2)f^y$ respectively. The maximum value of x is 10 and the maximum value of y is 14.

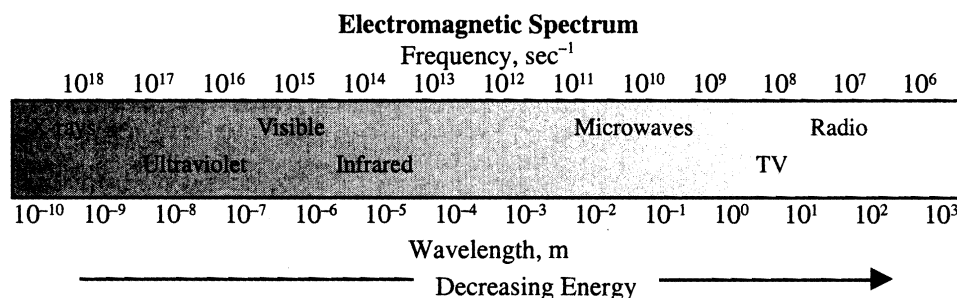
AS-7. The energy of a photon is greatest in the case of

- | | |
|--------------------|---------------------------|
| (A) X-rays. | (B) ultraviolet radiation |
| (C) visible light. | (D) infrared radiation. |

Atomic Structure

Knowledge Required: (1) The meaning of the term “photon.” (2) The names associated with regions of the electromagnetic spectrum. (3) The qualitative knowledge of the relationships between energy, frequency, and wavelength ($E = h\nu = \frac{hc}{\lambda}$).

Thinking it Through: A photon is a packet of energy, the quantity of which is related to the frequency of the electromagnetic radiation. Higher frequencies (ν) and shorter wavelengths (λ) correlate with higher photon energies. All electromagnetic radiation travels at the velocity of light (c). The labels associated with regions of the electromagnetic spectrum are shown in the schematic diagram. Visible light and infrared radiation, choices (C) and (D), are relatively low energy. Ultraviolet radiation, choice (B), is higher in energy, but X-rays, choice (A), are highest.



AS-8. Which ground-state electron configuration is possible for an atom in the second period?

- (A) $1s^2 2s^1 2p^1$ (B) $1s^2 2s^2 2d^1$ (C) $1s^2 2s^1$ (D) $1s^2 2p^4$

Knowledge Required: (1) The symbolic representation of electron configuration. (2) The meaning of the term “second period.” (3) The allowed values for principal and angular momentum quantum numbers. (4) The order of filling of electron orbitals. (5) The ability to correlate symbolic representations with positions in the periodic table.

Thinking it Through: For the second period, the principal quantum number, n , is equal to 2. For a principal quantum number of 2, the angular momentum (shape) quantum number, l , has allowed values of 0 or 1. The letter s is associated with $l = 0$, and the letter p is associated with $l = 1$. Choice (B) cannot be correct, since only s and p orbitals are possible when $n = 2$. Choice (D) is eliminated because the $2s$ orbital must fill before the $2p$ orbital. The $2s$ orbital has room for two electrons, which must be occupied before the $2p$ orbital begins to fill. Choice (A) is eliminated because the order of filling is incorrect. Choice (C) has the correct orbitals present for a second period element, and also displays the correct order for filling orbitals. It is the electron configuration for lithium.

AS-9. The number of unpaired electrons in a gaseous selenium atom is

- (A) 2 (B) 3 (C) 4 (D) 5

Atomic Structure

4. The number of neutrons in the $^{37}_{17}\text{Cl}^-$ ion is
(A) 17 (B) 20 (C) 21 (D) 37
5. An argon atom is isoelectronic with
(A) Cl (B) Ca (C) Ti^{4+} (D) N^{3-}
6. Which pair of particles has the same number of electrons?
(A) F^- , Mg^{2+} (B) Ne, Ar
(C) Br^- , Se (D) Al^{3+} , P^{3-}
7. Which ion has twenty-six electrons?
(A) Cr^{2+} (B) Fe^{2+} (C) Ni^{2+} (D) Cu^{2+}
8. Which statement is true?
(A) The nucleus of an atom contains neutrons and electrons.
(B) The atomic number of an element is the number of protons in one atom.
(C) The mass number of an atom is the number of protons in the nucleus plus the number of electrons outside.
(D) The number of electrons outside the nucleus is the same as the number of neutrons in the nucleus.
9. In what respect does an atom of magnesium, Mg, differ from a magnesium ion, Mg^{2+} ?
(A) The ion has an inert gas electron configuration; the atom does not.
(B) The positive charge on the nucleus of the ion is two units greater than the nuclear charge on the atom.
(C) The ion has two more protons than the atom.
(D) The ion has two more planetary electrons than the atom.
10. A sodium ion differs from a sodium atom in that the sodium ion
(A) has fewer electrons.
(B) is an isotope of sodium.
(C) exists only in solution.
(D) has a negative charge on its nucleus.
11. Which term best characterizes the relation of hydrogen to deuterium?
(A) allotropes (B) isomers (C) isotopes (D) polymers
12. The element X occurs naturally to the extent of 20.0% ^{12}X and 80.0% ^{13}X . The atomic mass of X is nearest
(A) 12.2 (B) 12.5 (C) 12.8 (D) 13.0

13. In which pair are the two species *both* isoelectronic and isotopic?
- (A) $^{40}_{20}\text{Ca}^{2+}$ and $^{40}_{18}\text{Ar}$ (B) $^{39}_{19}\text{K}^{+}$ and $^{40}_{19}\text{K}^{+}$
 (C) $^{24}_{12}\text{Mg}^{2+}$ and $^{25}_{12}\text{Mg}$ (D) $^{56}_{26}\text{Fe}^{2+}$ and $^{57}_{26}\text{Fe}^{3+}$
14. An atom of the element of atomic number 84 and mass number 199 emits an alpha particle. The residual atom after this change has an atomic number of ____ and a mass number of ____.
- (A) 82, 195 (B) 84, 203 (C) 85, 195 (D) 86, 199
15. In effecting nuclear changes by bombarding target nuclei with positively charged alpha particles, it is necessary to accelerate these particles to high speed because it is necessary to
- (A) drive the positive particles through the electron cloud.
 (B) overcome the force of repulsion of the nucleus.
 (C) focus the bombarding particle more accurately.
 (D) strip electrons from the atom.
16. What significant information about atomic structure came from the Millikan experiment using charged oil drops?
- (A) Millikan showed that cathode rays were identical to a stream of electrons coming from an atom.
 (B) Millikan confirmed that the neutron and proton were of about the same mass.
 (C) Millikan determined the magnitude of the charge on an electron.
 (D) Millikan proved that the mass of an atom was concentrated in the nucleus.
17. What is the valence electron configuration for the element in Period 5, Group 3A?
- (A) $5s^25p^1$ (B) $3s^23p^5$ (C) $3s^23p^3$ (D) $5s^25p^3$
18. Which electron configuration is *impossible*?
- (A) $1s^22s^22p^63s^2$ (B) $1s^22s^22p^62d^2$
 (C) $1s^22s^22p^63s^23p^6$ (D) $1s^22s^22p^53s^1$
19. The ground-state electronic configuration of the manganese atom, Mn, is
- (A) $1s^22s^22p^63s^23p^64s^24d^5$ (B) $1s^22s^22p^63s^23p^63d^7$
 (C) $1s^22s^22p^63s^23p^64s^24p^5$ (D) $1s^22s^22p^63s^23p^63d^54s^2$
20. Which species has this ground-state electron arrangement?
- $$1s^22s^22p^63s^23p^63d^{10}$$
- (A) Ni (B) Ni^{2+} (C) Zn (D) Zn^{2+}
21. The maximum number of electrons that can occupy an orbital labeled d_{xy} is
- (A) 1 (B) 2 (C) 3 (D) 4

Atomic Structure

22. An atom of Fe has two $4s$ electrons and six $3d$ electrons. How many *unpaired* electrons would there be in the Fe^{2+} ion?
- (A) one (B) two (C) three (D) four
23. Which of these species (is/are) paramagnetic?
- Ti^{4+} Fe^{2+} Zn^0
- (A) Fe^{2+} only (B) Zn^0 only
(C) Ti^{4+} and Fe^{2+} only (D) Fe^{2+} and Zn^0 only
24. Which of these electron diagrams could represent the ground state of the p valence electrons of carbon?
- (A) $\uparrow\downarrow$ — — (B) $\uparrow$ $\downarrow$ — —
(C) $\uparrow\uparrow$ — — (D) $\uparrow$ $\uparrow$ — —
25. The existence of discrete (quantized) energy levels in an atom may be inferred from
- (A) experiments on the photoelectric effect.
(B) diffraction of electrons by crystals.
(C) X-ray diffraction by crystals.
(D) atomic line spectra.
26. Which emission line in the hydrogen spectrum occurs at the highest frequency?
- (A) $n = 3 \rightarrow n = 1$ (B) $n = 4 \rightarrow n = 2$
(C) $n = 7 \rightarrow n = 5$ (D) $n = 10 \rightarrow n = 8$
27. Which set of quantum numbers is correct and consistent with $n = 4$?
- (A) $l = 3$ $m_l = -3$ $m_s = +\frac{1}{2}$ (B) $l = 4$ $m_l = +2$ $m_s = -\frac{1}{2}$
(C) $l = 2$ $m_l = +3$ $m_s = +\frac{1}{2}$ (D) $l = 3$ $m_l = -3$ $m_s = +1$
28. When an atom of an electropositive atom becomes an ion it
- (A) gains electrons. (B) becomes larger.
(C) emits an alpha particle. (D) does none of these.
29. The orbitals of $2p$ electrons are often represented as being
- (A) elliptical. (B) tetrahedral.
(C) dumbbell shaped. (D) spherical.
30. Helium, ${}^4_2\text{He}$, has two electrons in the $1s$ orbital. When it becomes singly ionized, forming He^+ ,
- (A) its spectrum resembles that of the hydrogen spectrum.
(B) the remaining electron is easier to remove.
(C) the nuclear charge has decreased by one.
(D) it has lost one atomic mass unit.

Answers to Study Questions

5. C
6. D
7. A
8. C

9. A
10. A

Answers to Practice Questions

11. C
12. C
13. B
14. A
15. B
16. C
17. A
18. B
19. D
20. D

21. B
22. D
23. A
24. D
25. D
26. A
27. A
28. D
29. C
30. A

Molecular Structure and Bonding

Understanding molecular structure requires knowledge of the spatial arrangement of atoms in a molecule. The geometric arrangement of atoms in molecules can be determined by experiment, but bonding theories and properties such as dipole moment, which is a measure of molecular polarity, can help predict and rationalize molecular shape and chemical behavior. This section of most general chemistry courses deals with valence-shell electron-pair repulsion (VSEPR) models, prediction of bond angles, valence bond theory, hybrid orbitals, and molecular orbital theory.

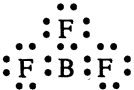
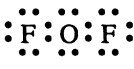
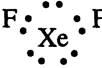
Study Questions

MS-1. Which molecule has exactly two unshared (lone) pairs of electrons on the central atom?

- (A) BF_3 (B) OF_2 (C) NF_3 (D) XeF_2

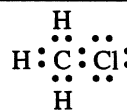
Knowledge Required: (1) The rules for drawing Lewis structures. (2) The exceptions to the octet rule.

Thinking it Through: Start by finding the total number of valence electrons that are available. The group number for the main group or representative (A group) elements gives this number. Fluorine, for example, is a member of Group VIIA and has 7 valence electrons. Distribute the electrons to satisfy the octet rule, keeping in mind the common exceptions.

BF_3	$3 + 3(7) = 24$ valence electrons		No unshared pairs on boron.
OF_2	$6 + 2(7) = 20$ valence electrons		Exactly two unshared pairs on oxygen.
NF_3	$5 + 3(7) = 26$ valence electrons		One unshared pair on nitrogen.
XeF_2	$8 + 2(7) = 22$ valence electrons		Three unshared pairs on xenon.

Note that there are two common exceptions here. Boron has too few valence electrons to form a full octet, and xenon can form an expanded octet, making both choices (A) and (D) incorrect. NF_3 has only one lone pair, so choice (C) is also incorrect. Only OF_2 has exactly two unshared pairs of electrons on the central atom, making choice (B) the correct answer.

MS-2. Consider the Lewis structure for CH_3Cl . What is the best description of the molecular shape?



- (A) bent (B) square
(C) square pyramidal (D) tetrahedral

Molecular Structure and Bonding

Knowledge Required: (1) The interpretation of Lewis structures. (2) The valence-shell-electron-pair-repulsion (VSEPR) model for predicting molecular shape.

Thinking it Through: The first step in applying the VSEPR method is to count the number of electron regions around the central atom. In this case, the Lewis structure is given, showing four pairs of electrons around the carbon atom. These electron pairs are all used to form bonds, and there are no lone pairs. Electron pair repulsion is minimized when the four electron pairs form a tetrahedron around the carbon atom, choice (D). Observe the similarity to methane, CH_4 , which was probably the first tetrahedral molecule you studied.

Note: In other questions you may be asked about the *tetrahedral bond angle*, which is 109.5° .

MS-3. The fact that BCl_3 is a planar molecule while NCl_3 is pyramidal can be explained several different ways. Which statement is the best rationalization?

- (A) Nitrogen is more electronegative than boron.
- (B) The nitrogen atom in NCl_3 has a lone pair of electrons, whereas the boron atom in BCl_3 does not.
- (C) The nitrogen atom is smaller than the boron atom.
- (D) The boron atom in BCl_3 is sp^3 hybridized, while the nitrogen atom in NCl_3 is sp^2 hybridized.

Knowledge Required: (1) The valence-shell-electron-pair-repulsion (VSEPR) model for predicting molecular shape. (2) The valence-bond theory for predicting hybridization of orbitals.

Thinking it Through: Choices (A) and (C) are both true statements, but neither provides the *reason* for the difference in geometries. The VSEPR model does provide a rationalization for the difference. Choice (D) has the hybridization reversed. NCl_3 has an sp^3 hybridized nitrogen atom, and BCl_3 has an sp^2 hybridized boron atom. Choice (B), the correct answer, recognizes that boron has only enough valence electrons to form three bonds, with no lone pairs. The six electrons arrange themselves to have 120° $\text{Cl}-\text{B}-\text{Cl}$ bond angles. On the other hand, when NCl_3 forms three bonds it has an electron pair left over to complete the octet. The four electron pairs arrange themselves in the shape of a tetrahedron around the nitrogen atom. Here are the Lewis structures and their three-dimensional representations:



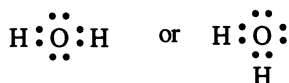
MS-4. According to the VSEPR model, the geometric structure of H_2O is

- (A) bent at an angle of 104.5° because of the greater repulsion of two lone pairs relative to that of the two bonding pairs.
- (B) bent at an angle of 120° because of the mutual repulsion of the six valence electrons on oxygen.
- (C) bent at an angle of 90° because of the perpendicular relationship of the oxygen p orbitals relative to each other.
- (D) linear with a bond angle of 180° because of the mutual repulsion of the bonding pairs of electrons.

Molecular Structure and Bonding

Knowledge Required: (1) Valence-shell-electron-pair-repulsion (VSEPR) model for predicting molecular shape. (2) Effect of lone pairs on predicted bond angles.

Thinking it Through: Recall or write the Lewis structure for the water molecule.



A Lewis structure by itself does not provide information about geometry. Use VSEPR to predict geometry by counting bonded and nonbonded electron pairs. If there are four bonded pairs to identical atoms, the exact tetrahedral angle of 109.5° can be expected. That is the case with CH_4 . In NH_3 , there are three bonded pairs and one nonbonded pair, so the bond angle closes slightly (to 107°) as the result of the lone pair requiring more space than the bonded pairs. In the water molecule, with two nonbonded pairs, the $\text{H}-\text{O}-\text{H}$ bond angle closes to 104.5° because of repulsion from the lone pairs. (A) is the correct choice. Choice (B) has the bond angle larger, rather than smaller, than the tetrahedral angle. Choice (C) treats the orbitals as if they had the geometry of atomic orbitals. Choice (D) is used to recognize the three-dimensional nature of molecules.

5. Sulfur dioxide can be described by the structures shown. This symbolism indicates that the



- (A) two bonds in SO_2 are of equal length, and the electronic distribution in the two sulfur-to-oxygen bonds is identical.
- (B) single bond is longer than the double bond, and the electronic distribution in the two sulfur-to-oxygen bonds differs.
- (C) electron pair in the SO_2 molecule alternates back and forth between the two sulfur-oxygen electron pairs, so that the two different bonds seem to exchange positions.
- (D) SO_2 molecule revolves, so that the two different sulfur-to-oxygen bonds seem to exchange positions.

Knowledge Required: (1) The meaning of resonance structures. (2) The interpretation of Lewis structures.

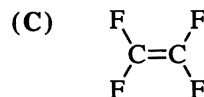
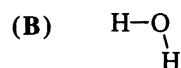
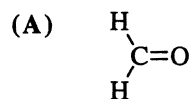
Thinking it Through: A single Lewis electron-dot formula cannot always adequately describe bonding between atoms. If two (or more) formulas satisfy the octet rule, the Lewis structures are both written and resonance is indicated with a double-headed arrow. This does *not* mean that the molecule oscillates between the structures. Rather, it means that the actual structure is intermediate between (or among) the resonance forms. Choices (C) and (D) suggest the largest movement of electrons or atoms, which does not occur. Choice (B) suggests that the sulfur-oxygen bonds are not equivalent, but they are. Correct choice (A) recognizes that the two bonds are identical. The need to write two forms reflects a limitation of the Lewis-dot model.

A compound consisting of an element having a low ionization energy and a second element having a high electron affinity is likely to have

- (A) covalent bonds.
- (B) metallic bonds.
- (C) coordinate covalent bonds.
- (D) ionic bonds.

Molecular Structure and Bonding

MS-9. Which of these is a nonpolar molecule?



Knowledge Required: (1) The relative electronegativity of atoms. (2) The properties of polar bonds. (3) The molecular geometry of molecules. (4) The relationship of dipole moment to molecular geometry.

Thinking it Through: For a molecule to be nonpolar when it has polar bonds, it must be symmetric enough that all the bond dipoles cancel each other out. For choice (A), the pull of the oxygen on the electrons shared with carbon is not cancelled from the other side, and the molecule is polar. The same is true for choices (B) and (D). Choice (C) contains polar carbon-fluorine bonds, with the polarity of each one cancelled by a carbon-fluorine bond on the opposite side of the molecule.

MS-10. For which molecule can the bonding be described in terms of sp^3 hybrid orbitals of the central atom?



Knowledge Required: (1) Geometry of hybrid orbitals. (2) Number of valence electrons for several atoms.

Thinking it Through: To form sp^3 hybrid orbitals, valence-bond theory uses one s orbital and three p orbitals to produce four equivalent orbitals oriented toward the corners of a tetrahedron. Each of the hybrid orbitals has a capacity of two electrons, for a maximum of eight electrons in the valence shell. When sulfur forms two-electron bonds with each of six fluorine atoms, as in choice (A), the sulfur must form a hybrid that uses *six* atomic orbitals rather than four. Similarly, the phosphorus atom in choice (C) must form hybrids that use *five* atomic orbitals. Boron, with only three valence electrons, cannot place one in each of four hybrid orbitals. It is only able to form sp^2 hybrid orbitals, except when a bonding partner provides both of the shared electrons. Boron is a sp^2 hybrid in BF_3 , choice (B). A nitrogen atom, with five valence electrons, forms bonds with three hydrogen atoms. There are eight valence electrons, enough to fill four sp^3 hybrid orbitals, choice (D).

Practice Questions

- NH_3 (pyramidal geometry) reacts with BF_3 (planar geometry) to form the addition compound, H_3NBF_3 . What is the geometry around the nitrogen and boron centers in the addition compound?
 - Both centers are tetrahedral.
 - Nitrogen is tetrahedral; boron is linear.
 - Nitrogen is pyramidal; boron is planar.
 - Nitrogen is planar; boron is pyramidal.
- The molecule of the type ML_4 consists of four single bonds and no lone pairs. What structure is it expected to assume?
 - square planar
 - trigonal planar
 - trigonal pyramidal
 - tetrahedral

Molecular Structure and Bonding

3. The shape that most closely describes the NF_3 molecule is

- (A) octahedral. (B) trigonal planar.
(C) trigonal pyramidal. (D) tetrahedral.

4. In which pair are the molecules geometrically similar?

- (A) SO_2 and CO_2 (B) PH_3 and BF_3
(C) CO_2 and OF_2 (D) SO_2 and O_3

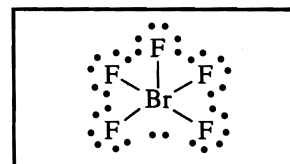
5. Which is planar?

- (A) NH_3 (B) SO_3^{2-} (C) CO_3^{2-} (D) CCl_4

6. Which is linear?

- (A) H_2S (B) NH_3 (C) NO_2 (D) CO_2

7. Consider the given Lewis structure for BrF_5 . What is the predicted shape for the molecule as a whole?



- (A) square pyramidal (B) trigonal bipyramidal
(C) trigonal pyramidal (D) octahedral

8. What is the shape of the XeF_4 molecule?

- (A) square planar (B) trigonal bipyramidal
(C) tetrahedral (D) trigonal pyramidal

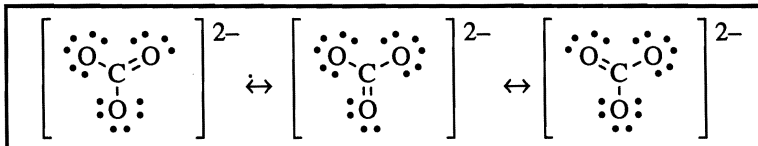
9. What set of species is arranged in order of increasing O-N-O bond angle?

- (A) NO_2^- , NO_2 , NO_2^+ (B) NO_2 , NO_2^- , NO_2^+
(C) NO_2^+ , NO_2 , NO_2^- (D) NO_2 , NO_2^+ , NO_2^-

10. Which has the largest bond angle?

- (A) angle O-S-O in SO_4^{2-} (B) angle Cl-C-Cl in HCCl_3
(C) angle F-Be-F in BeF_2 (D) angle H-O-H in H_2O

11. The structure of the CO_3^{2-} ion can be described in the Lewis formulation by these structures. This means that



- (A) two carbon-to-oxygen bonds are single bonds; the third is a double bond.
(B) three independent forms of the CO_3^{2-} ion coexist in equilibrium.
(C) the electrons must be rapidly exchanging among the three forms.
(D) the CO_3^{2-} ion exists in only one form: an average of the three principal structures shown.

Molecular Structure and Bonding

12. Which concept describes the formation of four equivalent, single, covalent bonds by carbon in its compounds that resemble methane, CH_4 ?
- (A) hydrogen bonding (B) hybridization
(C) sigma bonding (D) coordinate covalent bonding
13. Which type of hybrid orbital is used in CO_2 ?
- (A) sp (B) sp^2 (C) sp^3 (D) dsp^3
14. Which compound would be expected to have the largest dipole moment?
- (A) CO_2 (linear) (B) SO_2 (bent)
(C) BF_3 (trigonal planar) (D) CF_4 (tetrahedral)
15. The O-Si-O bond angles in SiO_2 (quartz) are closest to
- (A) 180° (B) 120°
(C) 110° (D) 100°
16. The molecule $\text{:}\ddot{\text{O}}=\text{C}=\ddot{\text{N}}-\text{H}$ has been detected in gas clouds between stars. The predicted C-N-H bond angle is about
- (A) 90° (B) 109° (C) 120° (D) 180°
17. Knowing that F is more electronegative than either B or P, what conclusion can be drawn from the fact that BF_3 has no dipole moment, but PF_3 does?
- (A) BF_3 is not spherically symmetrical, but PF_3 is.
(B) The BF_3 molecule must be trigonal planar.
(C) The BF_3 molecule must be linear.
(D) The atomic radius of P is larger than the atomic radius of B.
18. How many valence electrons are represented in the Lewis electron-dot structure for SO_2 ?
- (A) 6 (B) 8 (C) 18 (D) 32
19. Molecules of which compounds violate the octet rule?
- | | | | |
|---------------|--------------------------|----------------|----------------|
| NO_2 | CH_2Cl_2 | XeF_4 | NCl_3 |
| 1 | 2 | 3 | 4 |
- (A) 1 and 2 (B) 1 and 3 (C) 2 and 3 (D) 2 and 4
20. The boiling point of H_2O , compared with the other members of the series, can be explained by
- | Boiling Points of Group VIA Hydrides | | | |
|--------------------------------------|----------------------|-----------------------|-----------------------|
| H_2O | H_2S | H_2Se | H_2Te |
| 100°C | -61°C | -45°C | -2°C |
- (A) London dispersion forces. (B) dipole—induced dipole forces.
(C) hydrogen bonding. (D) nonpolar covalent bonding.

Molecular Structure and Bonding

21. The fact that $\text{Pt}(\text{NH}_3)_2\text{Cl}_2$ exists in two different isomeric forms offers evidence that the geometry is

- (A) octahedral. (B) square planar.
(C) tetrahedral. (D) trigonal planar.

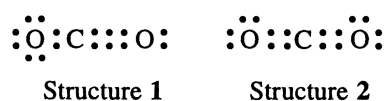
22. Which species has **both** covalent and ionic bonds?

- (A) NH_3BF_3 (B) H_3O^+ (C) NaKS (D) $\text{Mg}(\text{CN})_2$

23. When the carbon-carbon bonds in ethane (C_2H_6), ethene (C_2H_4), and benzene (C_6H_6), are arranged in order of increasing length (shortest bonds first), what is the correct order?

- (A) $\text{C}_2\text{H}_6 < \text{C}_2\text{H}_4 < \text{C}_6\text{H}_6$ (B) $\text{C}_2\text{H}_4 < \text{C}_2\text{H}_6 < \text{C}_6\text{H}_6$
(C) $\text{C}_6\text{H}_6 < \text{C}_2\text{H}_4 < \text{C}_2\text{H}_6$ (D) $\text{C}_2\text{H}_4 < \text{C}_6\text{H}_6 < \text{C}_2\text{H}_6$

24. These two electron-dot formulas for carbon dioxide both satisfy the octet rule, but one is preferred over the other. Which of the statements identifies the preferred structure and also the reason it is preferred?

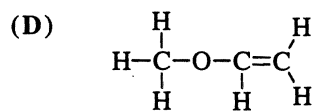
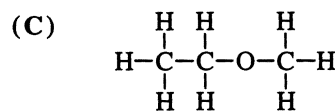
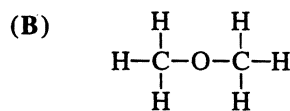
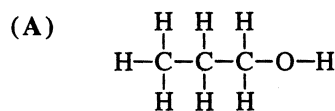


- (A) Structure 1 is preferred because the triple bond is stronger than the double bond.
(B) Structure 1 is preferred because the formal charge on carbon is negative.
(C) Structure 2 is preferred because the formal charge on each atom is zero.
(D) Structure 2 is preferred because the bonds are equal.

25. Which element is most likely to form a triple bond?

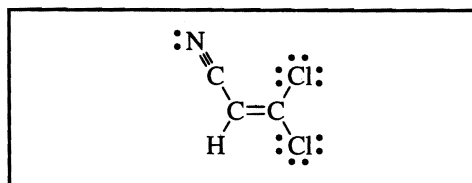
- (A) Pb (B) F (C) N (D) S

26. Which is an isomer of $\text{H}-\overset{\text{H}}{\underset{\text{H}}{\text{C}}}-\text{O}-\overset{\text{H}}{\underset{\text{H}}{\text{C}}}-\overset{\text{H}}{\underset{\text{H}}{\text{C}}}-\text{H}$?



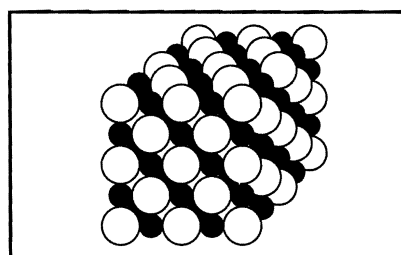
Molecular Structure and Bonding

27. How many sigma bonds and how many pi bonds are represented in this structure?



	Sigma Bonds	Pi Bonds
(A)	6	3
(B)	7	2
(C)	8	1
(D)	9	0

28. Sodium chloride, NaCl, crystallizes in a face-centered cubic lattice of chloride ions, with the smaller sodium ions occupying holes between the chloride ions. How many Cl^- ions are in contact with any single Na^+ ion?



- (A) 4 (B) 6 (C) 8 (D) 12
29. Resonance structures describe molecules that have
- (A) hybrid orbitals.
 (B) rapid equilibria.
 (C) resonating electrons.
 (D) multiple electron-dot formulas.

30. The complete Lewis structure of $\text{Cl}-\overset{\text{O}}{\underset{\text{Cl}}{\text{C}}}-\text{Cl}$ will have

- (A) at least one lone pair on each atom.
 (B) at least one double bond.
 (C) both polar and nonpolar bonds.
 (D) resonance forms.

Answers to Study Questions

- | | | |
|------|------|-------|
| 1. B | 5. A | 9. C |
| 2. D | 6. D | 10. D |
| 3. B | 7. C | |
| 4. A | 8. D | |

Answers to Practice Questions

- | | | |
|-------|-------|-------|
| 1. A | 11. D | 21. B |
| 2. D | 12. B | 22. D |
| 3. C | 13. A | 23. D |
| 4. D | 14. B | 24. C |
| 5. C | 15. C | 25. C |
| 6. D | 16. C | 26. A |
| 7. A | 17. B | 27. A |
| 8. A | 18. C | 28. B |
| 9. A | 19. B | 29. D |
| 10. C | 20. C | 30. B |

Multiple-Choice Questions

- A compound is found to consist of 34.9% sodium, 16.4% boron and 48.6% oxygen. What is its simplest formula?
(A) NaBO_2 (B) NaBO_3 (C) $\text{Na}_2\text{B}_4\text{O}_7$ (D) Na_3BO_3
- Upon analysis, a compound is found to contain 22.8% sodium, 21.8% boron, and 55.4% oxygen. What is its simplest formula?
(A) NaBO (B) NaB_2O_5 (C) $\text{Na}_2\text{B}_4\text{O}_7$ (D) Na_3BO_4
- A 4.08 g sample of a compound of nitrogen and oxygen contains 3.02 g of oxygen. What is the empirical formula?
(A) NO (B) NO_2 (C) N_2O (D) N_2O_5
- What is the percent by mass of oxygen in $\text{Ca}(\text{NO}_3)_2$?
(A) 29.3 % (B) 47.1 % (C) 58.5 % (D) 94.1 %

- Which of these compounds contains the greatest percentage of nitrogen?

Formula	Molar Mass
$\text{C}_6\text{H}_3\text{N}_3\text{O}_7$	$229. \text{ g} \cdot \text{mol}^{-1}$
$\text{CH}_4\text{N}_2\text{O}$	$60.1 \text{ g} \cdot \text{mol}^{-1}$
LiNH_2	$23.0 \text{ g} \cdot \text{mol}^{-1}$
$\text{Pb}(\text{N}_3)_2$	$291. \text{ g} \cdot \text{mol}^{-1}$

- (A) $\text{C}_6\text{H}_3\text{N}_3\text{O}_7$ (B) $\text{CH}_4\text{N}_2\text{O}$ (C) LiNH_2 (D) $\text{Pb}(\text{N}_3)_2$
- What mass of carbon is present in 0.500 mol of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$)?
(A) 60.0 g (B) 72.0 g (C) 90.0 g (D) 120. g
- How many atoms are in 1.50 g of Al?
(A) 0.0556 (B) 18.0 (C) 3.35×10^{22} (D) 2.44×10^{25}
- A sample of a compound of xenon and fluorine contains molecules of a single type; XeF_n , where n is a whole number. If 9.03×10^{20} of these XeF_n molecules have a mass of 0.311 g, what is the value of n ?
(A) 2 (B) 3 (C) 4 (D) 6

- What mass of carbon is present in 1.4×10^{20} molecules of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$)?

Formula	Molar Mass
$\text{C}_{12}\text{H}_{22}\text{O}_{11}$	$342 \text{ g} \cdot \text{mol}^{-1}$

- (A) $1.7 \times 10^{21} \text{ g}$ (B) $2.0 \times 10^{22} \text{ g}$ (C) $3.3 \times 10^{-2} \text{ g}$ (D) $2.8 \times 10^{-3} \text{ g}$
- A 2.000 g sample of an unknown metal, M, was completely burned in excess O_2 to yield 0.02224 mol of the metal oxide, M_2O_3 . What is the metal?
(A) Y (B) Ca (C) Al (D) Sc

Stoichiometry

11. A 3.41×10^{-6} g sample of a compound is known to contain 4.67×10^{16} molecules. This compound is
 (A) CO_2 (B) CH_4 (C) NH_3 (D) H_2O

12. Avogadro's number equals the number of

- (A) atoms in one mole of atoms. (B) molecules in one mole of O_2 .
 (C) marbles in one mole of marbles. (D) all of the above.

13. The number of atoms in 9.0 g of aluminum is the same as the number of atoms in

- (A) 8.1 g of magnesium. (B) 9.0 g of magnesium.
 (C) 12.1 g of magnesium. (D) 18.0 g of magnesium.

14. A single molecule of a certain compound has a mass of 3.4×10^{-22} g. Which value comes closest to the mass of a mole of this compound?

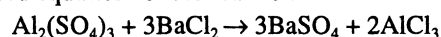
- (A) $50 \text{ g} \cdot \text{mol}^{-1}$ (B) $100 \text{ g} \cdot \text{mol}^{-1}$ (C) $150 \text{ g} \cdot \text{mol}^{-1}$ (D) $200 \text{ g} \cdot \text{mol}^{-1}$

15. What mass of KClO_3 will produce 48.0 g of oxygen gas, O_2 , if the decomposition of KClO_3 is complete?

Formula	Molar Mass
KClO_3	$122.5 \text{ g} \cdot \text{mol}^{-1}$

- (A) 61.3 g (B) 74.5 g (C) 122.5 g (D) 245.0 g

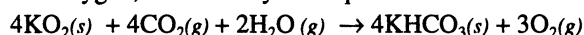
16. What is the maximum mass of aluminum chloride that could be obtained from 6.00 mol of barium chloride and excess aluminum sulfate? This is the balanced equation for the reaction.



Formula	Molar Mass
BaCl_2	$208.3 \text{ g} \cdot \text{mol}^{-1}$
AlCl_3	$133.3 \text{ g} \cdot \text{mol}^{-1}$

- (A) 1250 g (B) 801 g (C) 534 g (D) 134 g

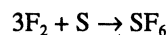
17. A self-contained breathing apparatus uses potassium superoxide, KO_2 , to convert the carbon dioxide and water in exhaled air into oxygen, as shown by this equation.



How many molecules of oxygen gas will be produced from the 0.0468 g of carbon dioxide that is exhaled in a typical breath?

- (A) 4.8×10^{20} (B) 6.4×10^{20} (C) 8.5×10^{20} (D) 1.9×10^{21}

18. A mixture containing 9 mol of F_2 and 4 mol of S is allowed to react. This equation represents the reaction that takes place.



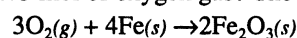
How many moles of F_2 remain after 3 mol of S have reacted?

- (A) 4 (B) 3 (C) 1 (D) 0

19. When 1.187 g of a metallic oxide is reduced with excess hydrogen, 1.054 g of the metal is produced. What is the metallic oxide?

- (A) Ag_2O (B) Cu_2O (C) K_2O (D) Ti_2O

20. How many moles of iron react with 1.75 mol of oxygen gas? The equation for the reaction is:



- (A) 1.31 mol (B) 1.75 mol (C) 2.33 mol (D) 5.25 mol

limiting reagent in a particular reaction can be recognized because it is the reagent that

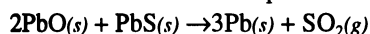
has the smallest coefficient in the balanced equation.

has the smallest mass in the reaction mixture.

is present in the smallest molar quantity.

would be used up first.

Consider this reaction used for the production of lead.

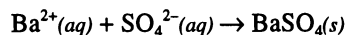


What is the maximum mass of lead that can be produced by the reaction of 57.33 g PbO and 33.80 g PbS?

Formula	Molar Mass
Pb	207.2 g·mol ⁻¹
PbO	223.2 g·mol ⁻¹
PbS	239.3 g·mol ⁻¹
SO ₂	64.07 g·mol ⁻¹

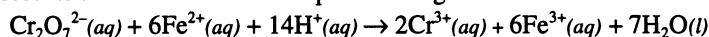
(A) 43.48 g (B) 72.75 g (C) 79.83 g (D) 87.80 g

What volume of 0.131 M BaCl₂ is required to react completely with 42.0 mL of 0.453 M Na₂SO₄? What is the net ionic equation for the reaction?



(A) 12.1 mL (B) 72.6 mL (C) 145 mL (D) 290 mL

In acidic solution, the dichromate ion, Cr₂O₇²⁻(aq) will oxidize Fe²⁺ to Fe³⁺ and form Cr³⁺. This net ionic equation represents the reaction that takes place during the reaction.



What volume of 0.100 M Cr₂O₇²⁻(aq) is required to oxidize 60.0 mL of 0.250 M Fe²⁺(aq)?

(A) 25.0 mL (B) 42.0 mL (C) 58.4 mL (D) 175 mL

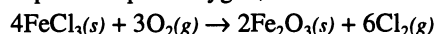
Consider this reaction. CaCO₃(s) → CaO(s) + CO₂(g)

What mass of CaCO₃ will produce 8.0 L of CO₂, measured at standard temperature and pressure conditions?

Formula	Molar Mass
CaCO ₃	100. g·mol ⁻¹

(A) 4.5 g (B) 12.5 g (C) 36 g (D) 280 g

When FeCl₃ is ignited in an atmosphere of pure oxygen, this reaction takes place.

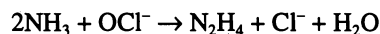


2.00 mol of FeCl₃ are ignited in the presence of 2.00 mol of O₂ gas, how much of which reagent is present in excess and therefore remains unreacted?

(A) 0.33 mol FeCl₃ remains unreacted (B) 0.67 mol FeCl₃ remains unreacted

(C) 0.25 mol O₂ remains unreacted (D) 0.50 mol O₂ remains unreacted

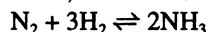
Ammonia, NH₃, reacts with the hypochlorite ion, OCl⁻, to produce hydrazine, N₂H₄. How many moles of hydrazine are produced from 5.85 mol of ammonia if the reaction has a 78.2% yield?



(A) 2.29 mol (B) 2.92 mol (C) 4.57 mol (D) 9.15 mol

Stoichiometry

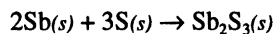
28. Ammonia can be prepared by the Haber process, shown in this equation.



If 2 mol of N_2 and 3 mol of H_2 are combined, the amount of NH_3 that would be formed if all of the limiting reactant were used up is known as the

- (A) limited yield. (B) percent yield. (C) product yield. (D) theoretical yield.

29. Antimony reacts with sulfur according to this equation.

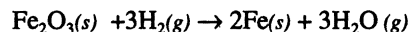


What is the percentage yield for a reaction in which 1.40 g of Sb_2S_3 is obtained from 1.73 g of antimony and a slight excess of sulfur?

- (A) 80.9% (B) 58.0% (C) 40.5% (D) 29.0%

Formula	Molar Mass
Sb_2S_3	$339.7 \text{ g}\cdot\text{mol}^{-1}$

30. When 4.50 g of Fe_2O_3 is reduced with excess H_2 in a furnace, 2.60 g of metallic iron is recovered. What is the percent yield? This is the equation representing the reaction.



- (A) 82.6% (B) 70.0% (C) 57.8% (D) 31.5%

Formula	Molar Mass
Fe_2O_3	$159.7 \text{ g}\cdot\text{mol}^{-1}$

Answers to Study Questions

- | | | |
|------|------|-------|
| 1. A | 5. B | 9. D |
| 2. D | 6. B | 10. B |
| 3. A | 7. C | |
| 4. B | 8. A | |

Answers to Practice Questions

- | | | |
|-------|-------|-------|
| 1. A | 11. A | 21. D |
| 2. C | 12. D | 22. C |
| 3. D | 13. A | 23. C |
| 4. C | 14. D | 24. A |
| 5. C | 15. C | 25. C |
| 6. B | 16. C | 26. A |
| 7. C | 17. A | 27. A |
| 8. C | 18. D | 28. D |
| 9. C | 19. B | 29. B |
| 10. D | 20. C | 30. A |

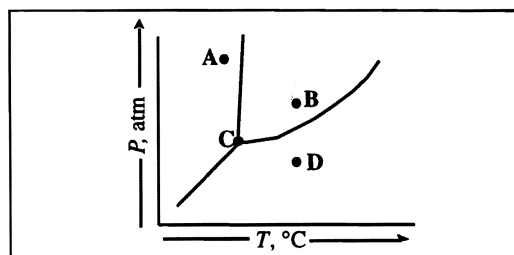
States of Matter/Solutions

Knowledge Required: (1) The meaning of specific heat, heat of melting, and melting point. (2) The interpretation of experimental data.

Thinking it Through: As a pure liquid cools undisturbed in a clean container, the temperature usually dips briefly below the melting point of the solid. When solid crystals begin to form, the temperature rises to the melting point of the solid (point A), where it remains until all the liquid has finally been converted to solid (point B). Neither A nor B are reasonable choices, because both liquid and solid are present from point A to point B. Choice (C) recognizes that the heat being removed from the liquid goes into changing the potential energy of the molecules. Kinetic energy remains the same. Choice (D) is an intensive property (temperature) rather than an extensive property (heat). The cooling curve interprets heat being removed from the liquid.

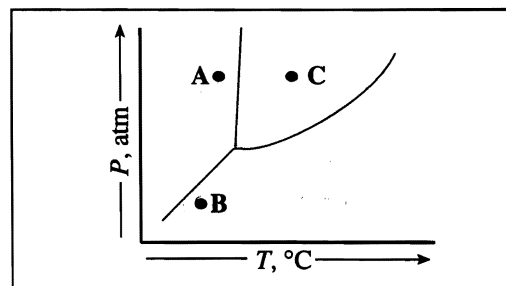
Practice Questions

1. In the phase diagram, which transition represents the condensation of a gas to a liquid?



- (A) B to A (B) D to B (C) C to D (D) A to D

2. Consider the phase diagram of a pure compound. Which statement applies?



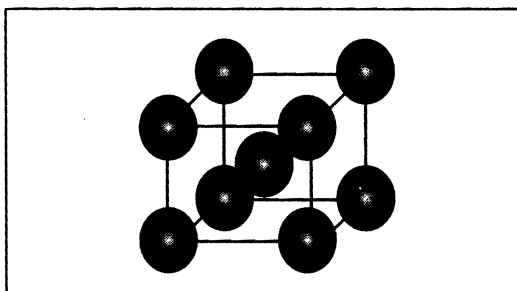
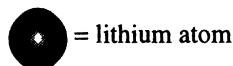
- (A) The path A $\rightarrow$ C represents sublimation.
(B) Following the path A $\rightarrow$ B $\rightarrow$ C, the compound would first liquefy and then vaporize.
(C) If the compound is in state A, continued reduction of the pressure (at constant temperature) will cause it to melt.
(D) None of these statements is correct.
3. A particular solid is soft, a poor conductor of heat and electricity, and has a low melting point. Generally, such a solid is classified as
- (A) ionic. (B) molecular.
(C) metallic. (D) network covalent.
4. Carbon dioxide, CO₂, in the form of dry ice would be classified as
- (A) an ionic solid. (B) a molecular solid.
(C) a polymeric solid. (D) a metallic solid.

5. The melting point of an impure compound is generally
- (A) higher than that of the pure solid.
 - (B) the same as that of the pure solid.
 - (C) lower than that of the pure solid.
 - (D) a function of the vapor pressure of the impurity.

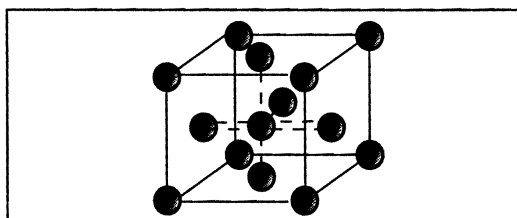
6. What is the vapor pressure of a solution with a benzene to octane molar ratio of 2:1?

Vapor Pressure at 50 °C	
benzene	280 mmHg
octane	400 mmHg

- (A) 120 mmHg (B) 320 mmHg (C) 400 mmHg (D) 680 mmHg
7. Lithium crystallizes in a body-centered cubic unit cell. What is the mass of one unit cell? Report your answer in grams.



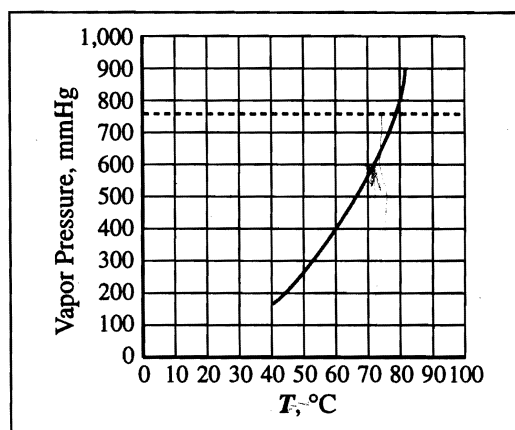
- (A) 1.15×10^{-23} g (B) 2.30×10^{-23} g (C) 5.75×10^{-23} g (D) 1.04×10^{-22} g
8. Copper crystallizes in a face-centered cubic lattice. If the edge of the unit cell is 362 pm, what is the radius of the copper atom?



- (A) 128 pm (B) 171 pm (C) 181 pm (D) 255 pm
9. Which factor affects the vapor pressure of a liquid?
- (A) temperature
 - (B) atmospheric pressure
 - (C) volume of the liquid
 - (D) surface area of the liquid
10. A gas or vapor may be liquefied only at temperatures
- (A) equal to the normal boiling point.
 - (B) above the normal boiling point.
 - (C) above the critical temperature.
 - (D) at or below the critical temperature.

States of Matter/Solutions

11. The graph shows how the vapor pressure of a liquid changes with temperature. Select the choice that best indicates the degree of correctness of this statement:
"The normal boiling point of the liquid is 78 °C."



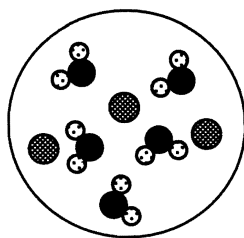
- (A) The statement is true.
 (B) The statement is probably true; additional data would be needed for a final decision.
 (C) The statement is probably false; additional data would be needed for a final decision.
 (D) The statement is false.
12. What is the molar mass of an ideal gas if a 0.622 g sample of this gas occupies a volume of 300. mL at 35 °C and 789 mm Hg?
- (A) 44.8 g·mol⁻¹ (B) 48.9 g·mol⁻¹ (C) 50.5 g·mol⁻¹ (D) 54.5 g·mol⁻¹
13. Which set of temperature and pressure conditions will cause a gas to exhibit the greatest deviation from ideal gas behavior?
- (A) 100 °C and 4 atm (B) 100 °C and 2 atm
 (C) -100 °C and 4 atm (D) 0 °C and 2 atm
14. A 500 mL gas sample is collected over water at a pressure of 740 mmHg and 25 °C. What is the volume of the dry gas at STP? (STP = 1 atm and 0 °C.)

Vapor Pressure of H ₂ O at 25 °C
24 mmHg

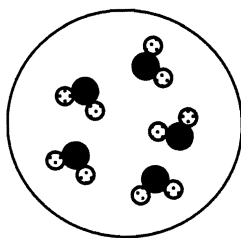
- (A) $500 \times \frac{740 + 24}{760} \times \frac{298}{273}$ (B) $500 \times \frac{740 - 24}{760} \times \frac{273}{298}$
 (C) $500 \times \frac{760}{740 + 24} \times \frac{273}{298}$ (D) $500 \times \frac{760}{740 - 24} \times \frac{298}{273}$

15. Which diagram represents the most concentrated solution?

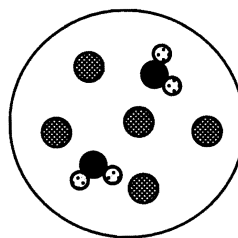
Note: ● represents a solute particle and ○ represents a water molecule.



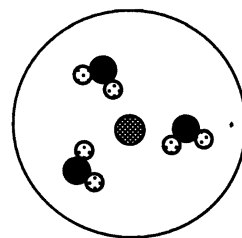
I



II



III



IV

(A) I

(B) II

(C) III

(D) IV

16. What is the molarity of a solution made by dissolving 8.56 g of sodium acetate in water and diluting to 750.0 mL?

Formula Molar Mass	
$\text{NaC}_2\text{H}_3\text{O}_2$	$82.03 \text{ g}\cdot\text{mol}^{-1}$

(A) 5.30 M (B) 0.139 M (C) 0.104 M (D) 0.0783 M

17. A student wants to prepare 250. mL of 0.10 M NaCl solution. Which procedure is most appropriate?

Formula Molar Mass	
NaCl	$58.4 \text{ g}\cdot\text{mol}^{-1}$

(A) Add 5.84 g of NaCl to 250. mL of H_2O .
 (B) Add 1.46 g of NaCl to 250. mL of H_2O .
 (C) Dissolve 5.84 g of NaCl in 50 mL of H_2O and dilute to 250. mL.
 (D) Dissolve 1.46 g of NaCl in 50 mL of H_2O and dilute to 250. mL

18. What volume of 12. M HCl solution is required to prepare exactly 500. mL of a 0.60 M HCl solution?

(A) 10. mL (B) 14 mL (C) 25 mL (D) 40. mL

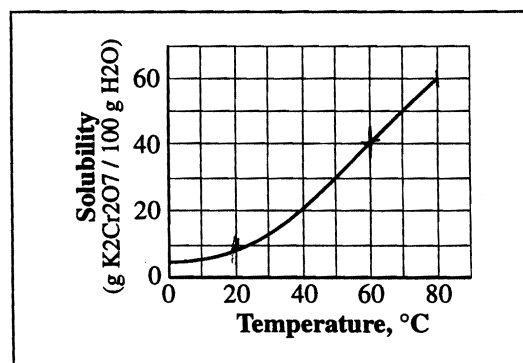
19. In which set are the substances arranged in order of decreasing solubility in water?

(A) $\text{Al}(\text{OH})_3 > \text{Mg}(\text{OH})_2 > \text{NaOH}$ (B) $\text{BaSO}_4 > \text{CaSO}_4 > \text{MgSO}_4$
 (C) $\text{CaCO}_3 > \text{NaHCO}_3 > \text{Na}_2\text{CO}_3$ (D) $\text{AgCl} > \text{AgBr} > \text{AgI}$

20. The solubility of a substance is 60 g per 100 mL of water at 15 °C. A solution of this substance is prepared by dissolving 75 g in 100 mL of water at 75 °C. The solution is then cooled slowly to 15 °C without any solid separating. The solution is

(A) supersaturated at 75 °C. (B) supersaturated at 15 °C.
 (C) unsaturated at 15 °C. (D) saturated at 15 °C.

21. A mixture of 100 g of $\text{K}_2\text{Cr}_2\text{O}_7$ and 200 g of water is stirred at 60 °C until no more of the salt dissolves. The resulting solution is poured off, leaving the undissolved solid behind. The solution is now cooled to 20 °C. What mass of $\text{K}_2\text{Cr}_2\text{O}_7$ crystallizes from the solution during the cooling?



(A) 24 g (B) 31 g (C) 43 g (D) 62 g

22. What mass of water is needed to dissolve 292.5 g of NaCl to produce a 0.25 m aqueous solution?

Formula Molar Mass	
NaCl	$58.5 \text{ g}\cdot\text{mol}^{-1}$

(A) 20 kg (B) 5.0 kg (C) 0.80 kg (D) 0.050 kg

23. What is the mole fraction of water in 200. g of 95% (by mass) ethanol, $\text{C}_2\text{H}_5\text{OH}$?

Formula Molar Mass	
$\text{C}_2\text{H}_5\text{OH}$	$46 \text{ g}\cdot\text{mol}^{-1}$

(A) 0.050 (B) 0.12 (C) 0.56 (D) 0.88

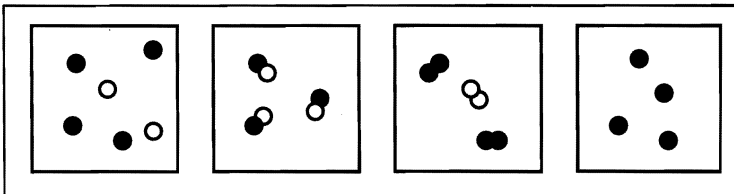
24. 800 g of ethanol, C_2H_5OH , was added to 8.0×10^3 g of water. How much would this lower the freezing point?

Freezing Point Depression Constant

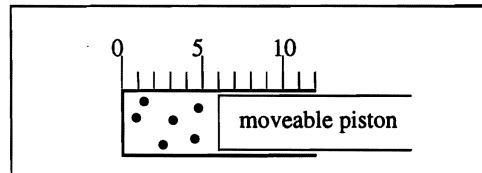
$$K_f \text{ for water} = 1.86 \text{ } ^\circ\text{C}\cdot m^{-1}$$

- (A) $3.2 \text{ } ^\circ\text{C}$ (B) $4.1 \text{ } ^\circ\text{C}$ (C) $8.2 \text{ } ^\circ\text{C}$ (D) $16 \text{ } ^\circ\text{C}$
25. The edge of a body-centered-cubic unit cell (which contains two atoms per unit cell) of an element Y was found to be 3.16×10^{-8} cm. The density of the metal is $19.35 \text{ g}\cdot\text{cm}^{-3}$. What is the approximate molar mass of Y?
- (A) $65.4 \text{ g}\cdot\text{mol}^{-1}$ (B) $92.0 \text{ g}\cdot\text{mol}^{-1}$ (C) $184 \text{ g}\cdot\text{mol}^{-1}$ (D) $238 \text{ g}\cdot\text{mol}^{-1}$
26. What is the number of nearest neighbors in a body-centered-cubic lattice?
- (A) 12 (B) 8 (C) 6 (D) 4
27. Carbon dioxide, CO_2 , in the form of dry ice would be classified as
- (A) an ionic solid. (B) a polymeric solid.
(C) a molecular solid. (D) a network solid.

28. Which diagrams represent pure substances?

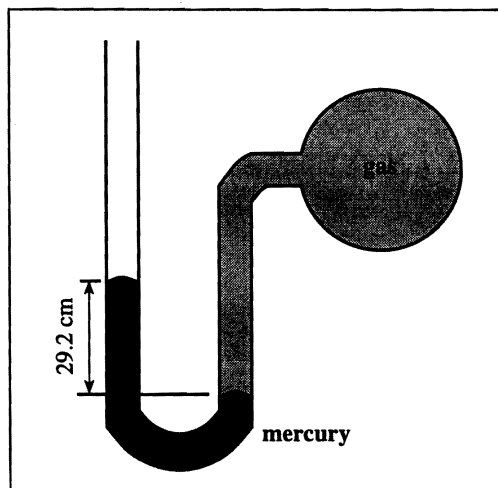


- (A) 1 and 2 (B) 1 and 3 (C) 2 and 3 (D) 2 and 4
29. Which diagram best represents the final system, if the pressure of the gas in this cylinder were doubled and the temperature increased from 200 K to 400 K?



- (A)
- (B)
- (C)
- (D)

30. An open-ended mercury manometer is used to measure the pressure exerted by a trapped gas as shown in the figure. Atmospheric pressure is 749 mmHg. What is the pressure of the trapped gas?



- | | |
|--------------|---------------|
| (A) 292 mmHg | (B) 457 mmHg |
| (C) 749 mmHg | (D) 1041 mmHg |

Answers to Study Questions

1. B
2. B
3. A
4. C

5. D
6. A
7. A
8. D

9. A
10. C

Answers to Practice Questions

1. B
2. D
3. B
4. B
5. C
6. B
7. B
8. A
9. A
10. D

11. A
12. C
13. C
14. B
15. C
16. B
17. D
18. C
19. D
20. B

21. D
22. A
23. B
24. B
25. C
26. B
27. C
28. D
29. C
30. D

Practice Questions

1. When a 45.0 g sample of an alloy at 100.0 °C is dropped into 100.0 g of water at 25.0 °C, the final temperature is 37.0 °C. What is the specific heat of the alloy?

Specific Heat Data	
H ₂ O(l)	4.184 J·g ⁻¹ ·°C ⁻¹

(A) 0.423 J·g⁻¹·°C⁻¹ (B) 1.77 J·g⁻¹·°C⁻¹ (C) 9.88 J·g⁻¹·°C⁻¹ (D) 48.8 J·g⁻¹·°C⁻¹

2. When 68.00 J of energy are added to a sample of gallium that is initially at 25.0 °C, the temperature rises to 38.0 °C. What is the volume of the sample?

Data for Gallium, Ga	
specific heat	0.372 J·g ⁻¹ ·°C ⁻¹
density	5.904 g·cm ⁻³

(A) 2.38 cm³ (B) 4.28 cm³ (C) 14.1 cm³ (D) 31.0 cm³

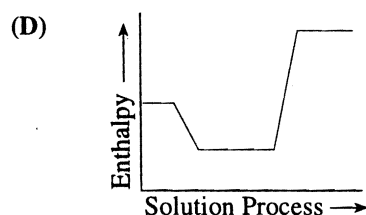
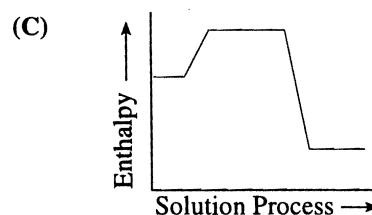
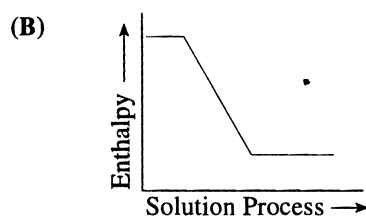
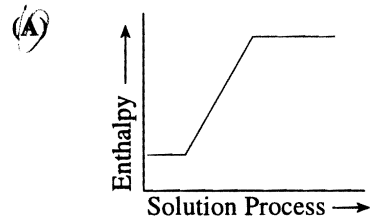
3. A student mixes 100 mL of 0.50 M NaOH with 100 mL of 0.50 M HCl in a Styrofoam® cup and observes a temperature increase of ΔT_1 . When she repeats this experiment using 200 mL of each solution, she observes a temperature change of ΔT_2 . If no heat is lost to the surroundings or absorbed by the Styrofoam cup, what is the relationship between ΔT_1 and ΔT_2 ?

(A) $\Delta T_2 = 4 \Delta T_1$ (B) $\Delta T_2 = 2 \Delta T_1$ (C) $\Delta T_2 = 0.5 \Delta T_1$ (D) $\Delta T_2 = \Delta T_1$

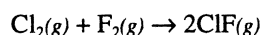
4. When a material in the liquid state is vaporized and then condensed to a liquid, the steps in the process are, respectively,

(A) exothermic and exothermic. (B) exothermic and endothermic.
(C) endothermic and exothermic. (D) endothermic and endothermic.

5. When Na₂S₂O₃·3H₂O dissolves in water, the solution gets cold. Which diagram best represents the change in enthalpy for the contents of the flask?



6. Calculate ΔH° for the chemical reaction



Bond Energies, kJ·mol ⁻¹	
F-F	159
Cl-Cl	243
Cl-F	255

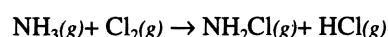
(A) -147 kJ (B) -108 kJ (C) +171 kJ (D) +912 kJ

Energetics

7. More heat is derived from cooling one gram of steam at 100 °C to water at 50 °C than from cooling one gram of liquid water at 100 °C to 50 °C because

(A) the steam is hotter than the water.
 (B) the steam occupies a greater volume than the water.
 (C) the density of water is greater than that of steam.
 (D) the heat of condensation is evolved.

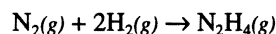
8. Calculate ΔH for this gas phase reaction.



Bond Energies, $\text{kJ}\cdot\text{mol}^{-1}$			
N-H	389	H-Cl	427
N-Cl	201	Cl-Cl	243

(A) -632 kJ (B) -188 kJ (C) +4 kJ (D) +431 kJ

9. Use the bond energies in the table to determine ΔH for the formation of hydrazine, N_2H_4 , from nitrogen and hydrogen according to this equation.



Bond Energies, $\text{kJ}\cdot\text{mol}^{-1}$			
N-N	159	H-H	436
N=N	201	H-N	389
N≡N	941		

(A) -711 kJ (B) -98 kJ (C) +98 kJ (D) +711 kJ

10. A 1.00 g sample of NH_4NO_3 is decomposed in a bomb calorimeter. The temperature increases by 6.12 °C. What is the molar heat of decomposition for ammonium nitrate?

Table of Data	
Molar mass, NH_4NO_3	80.0 $\text{g}\cdot\text{mol}^{-1}$
Calorimeter Constant	1.23 $\text{kJ}\cdot^\circ\text{C}^{-1}$

(A) -602 $\text{kJ}\cdot\text{mol}^{-1}$ (B) -398 $\text{kJ}\cdot\text{mol}^{-1}$ (C) 7.53 $\text{kJ}\cdot\text{mol}^{-1}$ (D) 164 $\text{kJ}\cdot\text{mol}^{-1}$

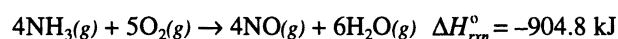
11. In a bomb calorimeter, reactions are carried out at

(A) constant pressure. (B) constant volume.
 (C) 1 atm pressure and 25 °C. (D) 1 atm pressure and 0 °C.

12. The standard enthalpy of formation, ΔH_f° , for nitrogen(IV) oxide, NO_2 , is the enthalpy change for which reaction?

(A) $\text{N}(g) + 2\text{O}(g) \rightarrow \text{NO}_2(g)$ (B) $\frac{1}{2}\text{N}_2(g) + \text{O}_2(g) \rightarrow \text{NO}_2(g)$
 (C) $\frac{1}{2}\text{N}_2\text{O}_4(g) \rightarrow \text{NO}_2(g)$ (D) $\text{NO}(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{NO}_2(g)$

13. The combustion of ammonia is represented by this equation.



What is the enthalpy of formation of $\text{NH}_3(g)$?

Enthalpy of Formation Data	
$\text{NO}(g)$	+90.4 $\text{kJ}\cdot\text{mol}^{-1}$
$\text{H}_2\text{O}(g)$	-241.8 $\text{kJ}\cdot\text{mol}^{-1}$

(A) -449 $\text{kJ}\cdot\text{mol}^{-1}$ (B) -46.1 $\text{kJ}\cdot\text{mol}^{-1}$ (C) -184 $\text{kJ}\cdot\text{mol}^{-1}$ (D) 227 $\text{kJ}\cdot\text{mol}^{-1}$

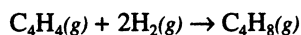
14. The standard enthalpy change for the formation of silver chloride from its elements is _____ than the enthalpy change for the formation of silver bromide.

Standard Enthalpies of Formation, 298 K

AgCl(s)	-127.1 kJ·mol ⁻¹
AgBr(s)	-100.4 kJ·mol ⁻¹

- (A) more exothermic
(B) less exothermic
(C) more endothermic
(D) less endothermic

15. What is the value of ΔH° for this reaction?

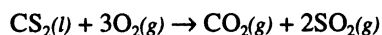


Enthalpies of Combustion

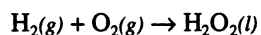
C ₄ H ₄ (g)	-2341 kJ·mol ⁻¹
H ₂ (g)	-286 kJ·mol ⁻¹
C ₄ H ₈ (g)	-2755 kJ·mol ⁻¹

- (A) -158 kJ (B) -128 kJ (C) +128 kJ (D) +158 kJ

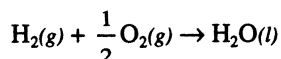
16. Given these values of ΔH° :



$$\Delta H^\circ = -1077 \text{ kJ}$$

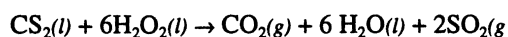


$$\Delta H^\circ = -188 \text{ kJ}$$



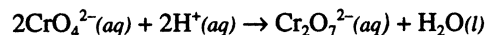
$$\Delta H^\circ = -286 \text{ kJ}$$

What is the value of ΔH° for this reaction?



- (A) -1175 kJ·mol⁻¹ (B) -1151 kJ·mol⁻¹ (C) -1665 kJ·mol⁻¹ (D) -3921 kJ·mol⁻¹

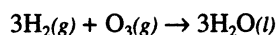
17. Use the standard enthalpies of formation in the table to calculate ΔH° for this reaction.

Substance ΔH_f° , kJ·mol⁻¹

CrO ₄ ²⁻ (aq)	-881.2
Cr ₂ O ₇ ²⁻ (aq)	-1490.3
H ⁺ (aq)	0
H ₂ O(l)	-285.8

- (A) +272.1 kJ (B) +13.7 kJ (C) -13.7 kJ (D) -272.1 kJ

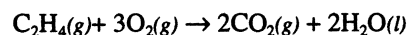
18. What is the value of ΔH° for this reaction?

Reaction ΔH , kJ

H ₂ (g) + 1/2O ₂ (g) → H ₂ O(l)	-286 kJ
3O ₂ (g) → 2O ₃ (g)	+271 kJ

- (A) -15 kJ (B) -558 kJ (C) -722 kJ (D) -994 kJ

19. Calculate the enthalpy of combustion of ethylene, C₂H₄, at 25 °C and one atmosphere pressure.

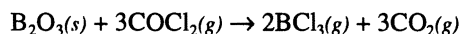
Compound ΔH_f° , kJ·mol⁻¹

C ₂ H ₄ (g)	+52.3
H ₂ O(l)	-285.8
CO ₂ (g)	-393.5

- (A) -1411 kJ (B) -1254 kJ (C) -732 kJ (D) -627 kJ

Energetics

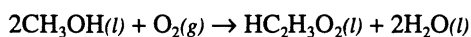
20. Use the given heats of formation to calculate the enthalpy change for this reaction.



Compound	$\Delta H_f^\circ, \text{kJ}\cdot\text{mol}^{-1}$
$\text{B}_2\text{O}_3(s)$	-1272.8
$\text{COCl}_2(g)$	-218.8
$\text{BCl}_3(g)$	-403.8
$\text{CO}_2(g)$	-393.5

- (A) 694.3 kJ (B) 354.9 kJ (C) -58.9 kJ (D) -3917.3 kJ

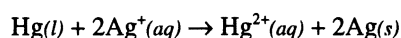
21. Using the given thermochemical data, what is ΔH° for this reaction?



Compound	$\Delta H_f^\circ, \text{kJ}\cdot\text{mol}^{-1}$
$\text{CH}_3\text{OH}(l)$	-238
$\text{HC}_2\text{H}_3\text{O}_2(l)$	-487
$\text{H}_2\text{O}(l)$	-286

- (A) 583 kJ (B) 535 kJ (C) -583 kJ (D) -535 kJ

22. What is the enthalpy change for this reaction?



Compound	$\Delta H_f^\circ, \text{kJ}\cdot\text{mol}^{-1}$
$\text{Ag}^+(aq)$	+105.6
$\text{Hg}^{2+}(aq)$	+171.1

- (A) +65.5 kJ (B) +40.1 kJ (C) -40.1 kJ (D) -65.5 kJ

23. For which of these processes is the value of ΔS expected to be negative?

- I. Sugar is dissolved in water.
 II. Steam is condensed.
 III. CaCO_3 is decomposed into CaO and CO_2 .

- (A) I only (B) I and III only
 (C) II only (D) II and III only

24. For which process is the entropy change per mole the largest at constant temperature?

- (A) $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$ (B) $\text{H}_2\text{O}(s) \rightarrow \text{H}_2\text{O}(g)$
 (C) $\text{H}_2\text{O}(s) \rightarrow \text{H}_2\text{O}(l)$ (D) $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(s)$

25. In which process is entropy *decreased*?

- (A) dissolving sugar in water (B) expanding a gas
 (C) evaporating a liquid (D) freezing water

26. When a liquid evaporates, which is true about the signs of the enthalpy and entropy changes?

- (A) ΔH is positive, ΔS is positive (B) ΔH is positive, ΔS is negative
 (C) ΔH is negative, ΔS is positive (D) ΔH is negative, ΔS is negative

27. A particular chemical reaction has a negative ΔH and a negative ΔS . Which statement is correct?

- (A) The reaction is spontaneous at all temperatures.
 (B) The reaction is nonspontaneous at all temperatures.
 (C) The reaction becomes spontaneous as temperature increases.
 (D) The reaction becomes spontaneous as temperature decreases.

28. When solid NH_4NO_3 is dissolved in water at 25°C , the temperature of the solution decreases. What is true about the signs of ΔH and ΔS for this process?

- (A) ΔH is negative, ΔS is positive (B) ΔH is negative, ΔS is negative
(C) ΔH is positive, ΔS is positive (D) ΔH is positive, ΔS is negative

29. For the process $\text{O}_2(\text{g}) \rightarrow 2\text{O}(\text{g})$, $\Delta H^\circ = +498 \text{ kJ}$. What would be predicted for the sign of ΔS_{rxn} and the conditions under which this reaction would be spontaneous?

	ΔS_{rxn}	Spontaneous
(A)	positive	at low temperatures only
(B)	positive	at high temperatures only
(C)	negative	at high temperatures only
(D)	negative	at low temperatures only

30. For which reaction, carried out at standard conditions, would both the enthalpy and entropy changes drive the reaction in the same direction?

- (A) $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$ $\Delta H = -571.1 \text{ kJ}$
(B) $2\text{Na}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{NaCl}(\text{s})$ $\Delta H = -822.0 \text{ kJ}$
(C) $\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$ $\Delta H = +67.7 \text{ kJ}$
(D) $2\text{NH}_3(\text{g}) \rightarrow \text{N}_2(\text{g}) + 3\text{H}_2(\text{g})$ $\Delta H = +92.4 \text{ kJ}$

Answers to Study Questions

- 1. B
- 2. A
- 3. A
- 4. A

- 5. D
- 6. B
- 7. C
- 8. D

- 9. B
- 10. A

Answers to Practice Questions

- 1. B
- 2. A
- 3. D
- 4. C
- 5. A
- 6. B
- 7. D
- 8. C
- 9. C
- 10. A

- 11. B
- 12. B
- 13. B
- 14. A
- 15. A
- 16. C
- 17. C
- 18. D
- 19. A
- 20. C

- 21. C
- 22. C
- 23. C
- 24. B
- 25. D
- 26. A
- 27. D
- 28. C
- 29. B
- 30. C

Dynamics

Knowledge Required: (1) How to use LeChatelier's principle. (2) The qualitative understanding of rate laws.

Thinking it Through: Considering LeChatelier's principle first, we see from the chemical equation that three moles of gas (2 mol SO₂ plus 1 mol of O₂) react to produce two moles of product. The equilibrium will shift to minimize the effect of the change, producing more SO₃. Choices (B) and (C) are therefore eliminated from consideration.

Increasing the pressure also increases the concentration of SO₂ and O₂, and one or the other (or both) are undoubtedly involved in the rate equation for this reaction. Increasing their concentrations is expected to increase the rate of reaction, leading to choice (A) as the correct answer. Choice (D) correctly recognizes the greater yield, but not the greater rate.

DY-10. In a chemical reaction involving the formation of an intermediate activated complex, which step must always be exothermic?

- | | | | |
|-----|-------------------------------|-----|------------------------------|
| (A) | reactants → products | (B) | products → reactants |
| (C) | reactants → activated complex | (D) | activated complex → products |

Knowledge Required: (1) The general understanding of multi-step reaction mechanisms. (2) The general shape of energy vs. reaction coordinate diagrams. (3) The meaning of the terms *exothermic* and *endothermic*.

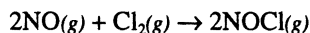
Thinking it Through: Spontaneous chemical reactions are usually exothermic, but not always. Choice (A) cannot be true. The reverse reaction of an exothermic reaction must be endothermic, so choice (B) is eliminated for the same reason as choice (A). The activated complex exists at the top of the energy barrier isolating reactants and products. Energetically speaking, it is downhill from the activated complex to either reactants or products. Choice (D), is the correct answer, is always exothermic. The remaining choice, (C), is always endothermic.

Practice Questions

1. The gas-phase reaction, $A_2 + B_2 \rightarrow 2AB$, proceeds by bimolecular collisions between A_2 and B_2 molecules. If the concentrations of both A_2 and B_2 are doubled, the reaction rate will change by a factor of

- (A) 1/2 (B) $\sqrt{2}$ (C) 2 (D) 4

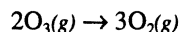
2. For the reaction of chlorine and nitric oxide,



doubling the concentration of chlorine doubles the rate of reaction. Doubling the concentration of both reactants increases the rate of reaction by a factor of eight. The reaction is

- (A) first order in both NO and Cl₂.
(B) first order in NO and second order in Cl₂.
(C) second order in NO and first order in Cl₂.
(D) second order in both NO and Cl₂.

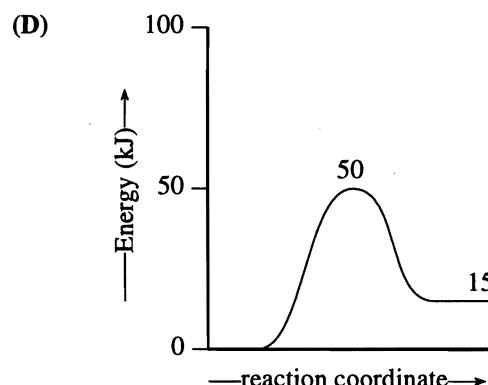
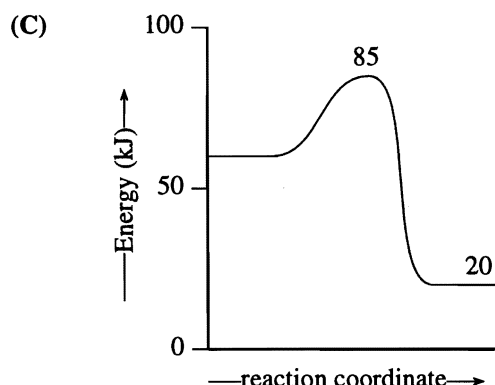
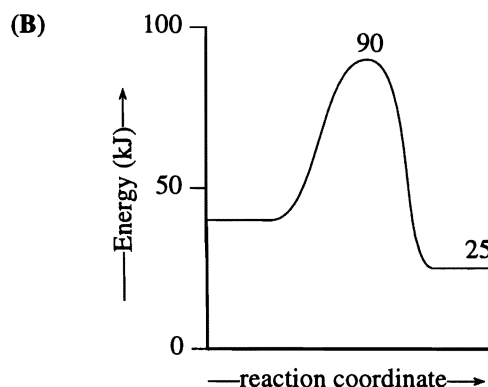
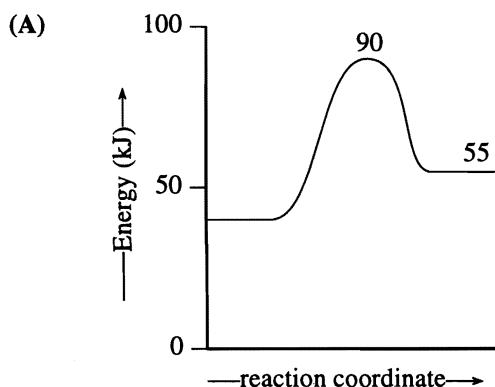
3. Under certain conditions, the average rate of *appearance* of oxygen gas in the reaction



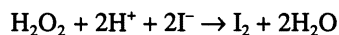
is $1.2 \times 10^{-3} \text{ atm}\cdot\text{s}^{-1}$. What is the average rate, expressed in $\text{atm}\cdot\text{s}^{-1}$, for the *disappearance* of O₃?

- (A) 8.0×10^{-4} (B) 1.2×10^{-3} (C) 1.8×10^{-3} (D) 5.3×10^{-3}

4. Which reaction coordinate diagram represents a reaction in which the activation energy, E_a , is $50 \text{ kJ}\cdot\text{mol}^{-1}$ and the ΔH_{rxn} is $-15 \text{ kJ}\cdot\text{mol}^{-1}$?



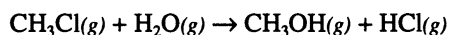
5. The rate law for the reaction



is $\text{rate} = k [\text{H}_2\text{O}_2] [\text{I}^-]$. The overall order of the reaction is

- (A) five. (B) three. (C) two. (D) one.
6. The activity of a radioisotope is 3000 counts per minute at one time and 2736 counts per minute 48 hours later. What is the half-life of the radioisotope?
- (A) 831 hr (B) 521 hr (C) 361 hr (D) 1.44 hr
7. The half-life for the first order conversion of cyclobutane to ethylene,
- $$\text{C}_4\text{H}_8(\text{g}) \rightarrow 2\text{C}_2\text{H}_4(\text{g})$$
- is 22.7 s at a particular temperature. How many seconds are needed for the partial pressure of cyclobutane to decrease from 100 mmHg to 10 mmHg?
- (A) 52.0 s (B) 75.4 s (C) 90.0 s (D) 227 s
8. A plot of reactant concentration as a function of time gives a straight line. What is the order of the reaction for this reactant?
- (A) zero (B) first (C) second (D) third

9. When the reaction

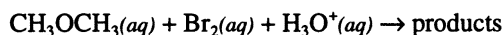


was studied, the tabulated data were obtained. Based on these data, what are the reaction orders?

Initial Concentrations, M; Initial Rates, $\text{M}\cdot\text{s}^{-1}$			
Exp	CH_3Cl	H_2O	Rate
1	0.100	0.100	0.182
2	0.200	0.200	1.45
3	0.200	0.400	5.81

- (A) CH_3Cl : first order H_2O : first order (B) CH_3Cl : first order H_2O : second order
 (C) CH_3Cl : second order H_2O : first order (D) CH_3Cl : second order H_2O : second order

10. The reaction between acetone and bromine in acid solution is represented by the equation



The tabulated kinetic data were gathered. Based on these data, the experimental rate law is

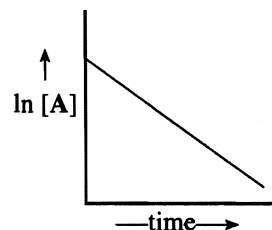
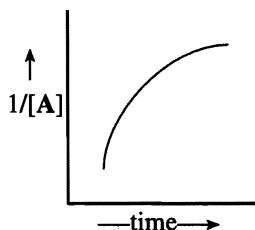
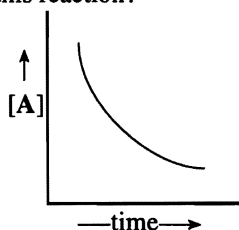
Initial Concentrations, M; Initial Rates, $\text{M}\cdot\text{s}^{-1}$				
Exp	CH_3OCH_3	Br_2	H_3O^+	Rate
1	0.30	0.050	0.050	5.8×10^{-5}
2	0.30	0.100	0.050	5.8×10^{-5}
3	0.30	0.050	0.100	1.2×10^{-4}
4	0.40	0.050	0.200	3.2×10^{-4}

- (A) $\text{rate} = k [\text{CH}_3\text{OCH}_3]^1 [\text{Br}_2]^1 [\text{H}_3\text{O}^+]^1$ (B) $\text{rate} = k [\text{CH}_3\text{OCH}_3]^1 [\text{Br}_2]^0 [\text{H}_3\text{O}^+]^1$
 (C) $\text{rate} = k [\text{CH}_3\text{OCH}_3]^0 [\text{Br}_2]^0 [\text{H}_3\text{O}^+]^2$ (D) $\text{rate} = k [\text{CH}_3\text{OCH}_3]^1 [\text{Br}_2]^1 [\text{H}_3\text{O}^+]^0$

11. All of these changes increase the value of the rate constant for a reaction
- except*

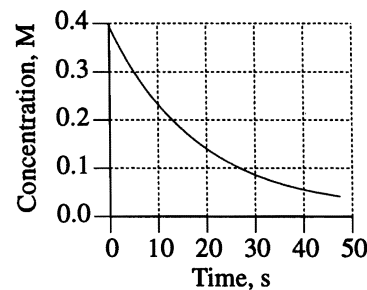
- (A) decreasing the activation energy.
 (B) raising the temperature.
 (C) adding a catalyst.
 (D) increasing the concentration of reactants.

12. The experimental data from a certain reaction gives these three graphs. What is the most likely order for this reaction?



- (A) zero (B) first (C) second (D) third

13. What is the initial rate of the reaction (in
- $\text{mol}\cdot\text{L}^{-1}\cdot\text{s}^{-1}$
-) depicted by this graph?



- (A) 0.02 (B) 0.01 (C) 0.008 (D) 0.005

1. Which function of $[X]$, plotted against time, will give a straight line for a second order reaction?
- (A) $[X]$ (B) $[X]^2$ (C) $\ln [X]$ (D) $1/[X]$

5. The Arrhenius equation describes the relationship between the rate constant, k , and the energy of activation, E_a .

$$k = Ae^{-E_a/RT}$$

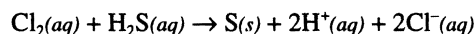
In this equation, A is an empirical constant, R is the ideal-gas constant, e is the base of natural logarithms, and T is the absolute temperature. According to the Arrhenius equation,

- (A) at constant temperature, reactions with lower activation energies proceed more rapidly.
(B) at constant temperature, reactions with lower activation energies proceed less rapidly.
(C) at constant energy of activation, reactions at lower temperatures proceed more rapidly.
(D) at constant energy of activation, reactions with smaller values of A proceed more rapidly.
16. What will be the effect of increasing the temperature of reactants that are known to undergo an endothermic reaction?
- (A) Both the rate of reaction and the value of the equilibrium constant increase.
(B) The rate of reaction increases and the value of the equilibrium constant decreases.
(C) The rate of reaction decreases and the value of the equilibrium constant increases.
(D) The rate of reaction increases and the value of the equilibrium constant is unchanged.
17. A change in temperature from 10°C to 20°C is found to double the rate of a particular chemical reaction. How did the change in temperature affect the reacting molecules?
- (A) The average velocity of the molecules doubled.
(B) The average energy of the molecules doubled.
(C) The number of collisions per second doubled.
(D) The number of molecules above the reaction energy threshold doubled.
18. The value for the rate constant of a reaction can generally be expected to
- (A) decrease with increasing temperature.
(B) increase with increasing temperature.
(C) decrease with increasing temperature only when the reaction is exothermic.
(D) increase with increasing temperature only when the reaction is exothermic.
19. Two reactions with different activation energies have the same rate at room temperature. Which statement correctly describes the rates of these two reactions at the same, higher temperature?
- (A) The reaction with the larger activation energy will be faster.
(B) The reaction with the smaller activation energy will be faster.
(C) The two reactions will continue to occur at the same rates.
(D) A prediction cannot be made without additional information.

20. A catalyst increases the rate of a reaction by

- (A) changing the mechanism of the reaction.
- (B) increasing the activation energy of the reaction.
- (C) increasing the concentration of one or more of the products.
- (D) decreasing the difference in relative energy of the reactants and products.

21. Consider the reaction



The rate equation for this reaction is

$$\text{rate} = k [\text{Cl}_2] [\text{H}_2\text{S}]$$

Which of these mechanisms is (or are) consistent with this rate equation?

- I** $\text{Cl}_2 + \text{H}_2\text{S} \rightarrow \text{H}^+ + \text{Cl}^- + \text{Cl}^+ + \text{HS}^-$ (slow)
 $\text{Cl}^+ + \text{HS}^- \rightarrow \text{H}^+ + \text{Cl}^- + \text{S}$ (fast)
- II** $\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}^-$ (fast equilibrium)
 $\text{Cl}_2 + \text{HS}^- \rightarrow 2\text{Cl}^- + \text{H}^+ + \text{S}$ (slow)
- (A) I only (B) II only
 (C) Both I and II (D) Neither I or II

22. A certain reaction has a $\Delta H = -75$ kJ and an activation energy of 40 kJ. A catalyst is found that lowers the activation energy of the forward reaction by 15 kJ. What is the activation energy of the reverse reaction in the presence of this same catalyst?

- (A) 25 kJ (B) 60 kJ (C) 90 kJ (D) 100 kJ

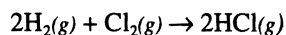
23. What is the relationship between the equilibrium constant (K_c) of a reaction and the rate constants for the forward (k_f) and reverse (k_r) reactions?

- (A) $K_c = k_f k_r$ (B) $K_c = k_f / k_r$
 (C) $K_c = 1 / (k_f k_r)$ (D) $K_c = k_f - k_r$

24. If the half-life of a reaction is independent of concentration, the reaction can be

- I first order II second order III zero order
- (A) I and II only. (B) II and III only.
 (C) I only. (D) II only.

25. Initial rate data for the reaction

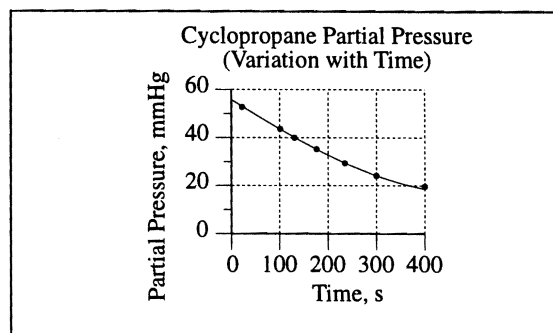


are given in the table. What is the rate law for this reaction?

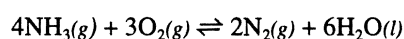
Experiment	$[\text{H}_2]_{\text{initial}}$	$[\text{Cl}_2]_{\text{initial}}$	Rate in $\text{M}\cdot\text{s}^{-1}$
1	0.0020 M	0.0050 M	2.5×10^{-3}
2	0.0020 M	0.0025 M	1.3×10^{-3}
3	0.0015 M	0.0025 M	1.3×10^{-3}
4	0.0050 M	0.0010 M	0.5×10^{-3}

- (A) $\text{rate} = k [\text{Cl}_2]^2$ (B) $\text{rate} = k [\text{Cl}_2]$
 (C) $\text{rate} = k [\text{H}_2]$ (D) $\text{rate} = k [\text{H}_2] [\text{Cl}_2]$

26. The rate of conversion from cyclopropane to propene in the gas phase at constant temperature is a first order reaction. From the information in the graph, which statement about the half-life of the conversion reaction is true?

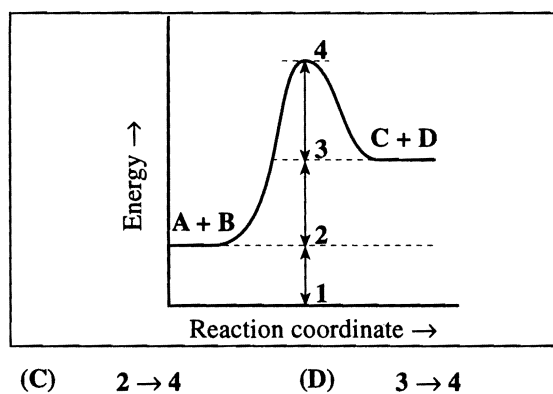


- (A) It is approximately equal to 150 s. (B) It is approximately equal to 270 s.
(C) It is approximately equal to 420 s. (D) It cannot be determined.
27. Consider this reaction.

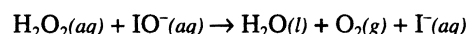
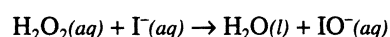


If the rate of formation of N_2 is $0.10 \text{ M}\cdot\text{s}^{-1}$, what is the corresponding rate of disappearance of O_2 ?

- (A) $1.5 \text{ M}\cdot\text{s}^{-1}$
(C) $0.15 \text{ M}\cdot\text{s}^{-1}$
28. Which line segment represents the activation energy for the reaction between C and D to form A and B?



- (A) $1 \rightarrow 4$ (B) $2 \rightarrow 3$ (C) $2 \rightarrow 4$ (D) $3 \rightarrow 4$
29. The decomposition of hydrogen peroxide in the presence of iodide ion is believed to occur via this mechanism.



In this mechanism, $\text{I}^-(aq)$ is

- (A) a catalyst. (B) a reactant in the overall reaction.
(C) the activated complex. (D) a product of the overall reaction.
30. Which statement most accurately describes the behavior of a catalyst?
- (A) A catalyst increases the ΔG of a reaction and hence the forward rate.
(B) A catalyst reduces the ΔH of a reaction and hence the temperature needed to produce products.
(C) A catalyst reduces the activation energy for a reaction and increases the rate of a reaction.
(D) A catalyst increases the equilibrium constant and final product concentrations.

Answers to Study Questions

- | | | |
|------|------|-------|
| 1. B | 5. B | 9. A |
| 2. B | 6. A | 10. D |
| 3. A | 7. C | |
| 4. D | 8. A | |

Answers to Practice Questions

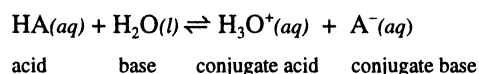
- | | | |
|-------|-------|-------|
| 1. D | 11. D | 21. A |
| 2. C | 12. B | 22. D |
| 3. A | 13. A | 23. B |
| 4. B | 14. D | 24. C |
| 5. C | 15. A | 25. B |
| 6. C | 16. A | 26. B |
| 7. B | 17. D | 27. C |
| 8. A | 18. B | 28. D |
| 9. B | 19. A | 29. A |
| 10. B | 20. A | 30. C |

Equilibrium Values, K_s

HC ₂ H ₃ O ₂	1.8×10^{-5}
HCN	6.2×10^{-10}
HF	7.2×10^{-4}

- (A) CN^- , $\text{C}_2\text{H}_3\text{O}_2^-$, F^- (B) $\text{C}_2\text{H}_3\text{O}_2^-$, CN^- , F^-
(C) F^- , $\text{C}_2\text{H}_3\text{O}_2^-$, CN^- (D) F^- , CN^- , $\text{C}_2\text{H}_3\text{O}_2^-$

Thinking it Through: Strong acids are completely ionized in water, so equilibrium constant expressions are not used for them. A weak acid is only partially ionized. The value of the equilibrium constant for a weak acid gives information about the degree of ionization.


$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$
$$\text{A}^{-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HA}(\text{aq}) + \text{OH}^{-}(\text{aq})$$

base acid conjugate acid conjugate base

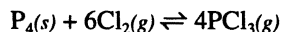
Choice (C) correctly lists the anions in their order of increasing base strength, based on decreasing acid strength. The only other choice that represents a common error is choice (A), which might indicate that you were choosing based on the acid strength, not the base strength. Both choices (B) and (D) are ineffective in this question, because they are based on nothing except possibly the order (or reverse order) of the acids as listed in the table of equilibrium values.

Practice Questions

1. Which is a proper description of chemical equilibrium?
- (A) The frequencies of reactant and of product collisions are identical.
- (B) The concentrations of products and reactants are identical.
- (C) The velocities of product and reactant molecules are identical.
- (D) Reactant molecules react to form products as fast as product molecules are reacting to form reactants.

Equilibrium

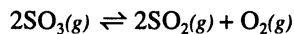
2. Consider this reaction.



What is the correct K_c expression for this reaction?

- (A) $K_c = \frac{1}{[\text{Cl}_2]^6}$ (B) $K_c = \frac{[\text{PCl}_3]^4}{[\text{P}_4][\text{Cl}_2]^6}$
 (C) $K_c = \frac{[\text{PCl}_3]^4}{[\text{Cl}_2]^6}$ (D) $K_c = \frac{[\text{P}_4][\text{Cl}_2]^6}{[\text{PCl}_3]^4}$

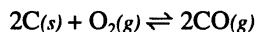
3. Consider this reaction.



What is the correct K_p expression for this reaction?

- (A) $K_p = \frac{P_{\text{SO}_2}^2 P_{\text{O}_2}}{P_{\text{SO}_3}^2}$ (B) $K_p = \frac{P_{\text{SO}_2} P_{\text{O}_2}}{P_{\text{SO}_3}}$
 (C) $K_p = \frac{(2P_{\text{SO}_2})^2 P_{\text{O}_2}}{(2P_{\text{SO}_3})^2}$ (D) $K_p = \frac{2P_{\text{SO}_2}^2 P_{\text{O}_2}}{2P_{\text{SO}_3}^2}$

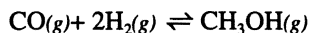
4. Consider this reaction.



What is the equilibrium expression for this reaction?

- (A) $K_c = \frac{[\text{CO}]}{[\text{C}][\text{O}_2]}$ (B) $K_c = \frac{[\text{CO}]^2}{[\text{C}]^2[\text{O}_2]}$
 (C) $K_c = \frac{[2\text{CO}]}{[2\text{C}][\text{O}_2]}$ (D) $K_c = \frac{[\text{CO}]^2}{[\text{O}_2]}$

5. Carbon monoxide gas reacts with hydrogen gas at elevated temperatures to form methanol according to this equation.



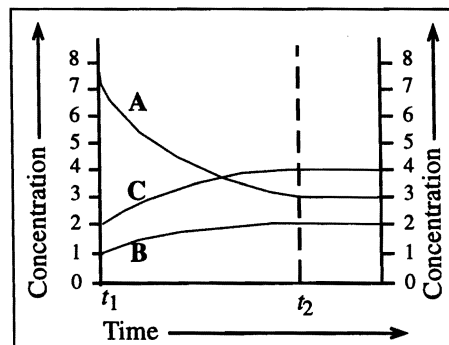
When 0.40 mol of CO and 0.30 mol of H_2 are allowed to reach equilibrium in a 1.0 L container, 0.060 mol of CH_3OH are formed. What is the value of K_c ?

- (A) 0.50 (B) 0.98 (C) 1.7 (D) 5.4

6. The graph shows the variation of concentration with time for this reaction at 25 °C.



What is the value of the equilibrium constant at time t_2 ?



- (A) 1.2 (B) 0.84 (C) 0.57 (D) 0.22

7. The value of an equilibrium constant can be used to predict each of these *except* the

- (A) direction of a reaction.
- (B) extent of a reaction.
- (C) quantity of reactant(s) remaining at equilibrium.
- (D) time required to reach equilibrium.

8. Chemical equilibrium is the result of

- (A) formation of products equal in mass to the mass of the reactants.
- (B) the unavailability of one of reactants.
- (C) a stoppage of further reaction.
- (D) opposing reactions attaining equal speeds.

9. Consider this gas-phase reaction.



Which reaction characteristics will be affected by a change in temperature?

1. value of equilibrium constant 2. equilibrium concentrations

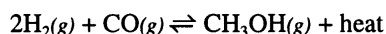
- (A) 1 only (B) 2 only (C) 1 and 2 (D) neither 1 nor 2

10. Which factors will affect both the position of equilibrium and the value of the equilibrium constant for this reaction?



- (A) increasing the volume of the container
- (B) adding more nitrogen gas
- (C) removing ammonia gas
- (D) lowering the temperature

11. The reversible reaction



is carried out by mixing carbon monoxide and hydrogen gases in a closed vessel under high pressure with a suitable catalyst. After equilibrium is established at high temperature and pressure, all three substances are present. If the pressure on the system is lowered, with the temperature kept constant, what will be the result?

- (A) The amount of CH_3OH will be increased.
- (B) The amount of CH_3OH will be decreased.
- (C) The amount of each substance will be unchanged.
- (D) The amount of each substance will be increased.

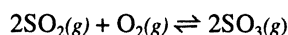
12. Consider this reaction, carried out at constant volume.



The concentration of $\text{O}_2(\text{g})$ at equilibrium increases if

- (A) SO_2 is added to the system.
- (B) SO_3 is added to the system.
- (C) the temperature of the system is lowered.
- (D) an inert gas is added to the system.

13. Consider this reaction, carried out at constant temperature and volume.



What is the effect of removing some SO_3 from a system initially at equilibrium?

- (A) The concentration of SO_2 decreases more than the concentration of O_2 .
- (B) The concentration of SO_2 increases more than the concentration of O_2 .
- (C) The concentration of SO_2 and O_2 remain the same.
- (D) The concentration of SO_2 and O_2 decrease equally.

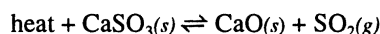
14. Consider this reaction, carried out at constant temperature and volume.



How can the position of equilibrium for this reaction be shifted to the right?

- (A) addition of a catalyst
- (B) removal of Cl_2
- (C) addition of an inert gas at constant volume
- (D) removal of PCl_5

15. Consider this reaction.



What change will cause an increase in the pressure of $\text{SO}_2(g)$ when equilibrium is re-established?

- | | |
|---|--|
| (A) increasing the reaction temperature | (B) decreasing the volume of the container |
| (C) adding some more CaSO_3 | (D) removing some of the $\text{CaO}(s)$ |

16. In which reaction would an increase in pressure at constant temperature have no effect on the relative amounts of the substances present in the equilibrium mixture? All substances are gases.

- | | |
|--|---|
| (A) $2\text{NO} + \text{O}_2 \rightleftharpoons 2\text{NO}_2 + \text{heat}$ | (B) $\text{heat} + \text{N}_2 + \text{O}_2 \rightleftharpoons 2\text{NO}$ |
| (C) $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3 + \text{heat}$ | (D) $2\text{CO} + \text{O}_2 \rightleftharpoons 2\text{CO}_2 + \text{heat}$ |

17. The numerical value of the equilibrium constant for any chemical change is affected by changing the

- | | |
|-----------------------------|------------------------------------|
| (A) nature of the catalyst. | (B) concentration of the products. |
| (C) the pressure. | (D) the temperature. |

18. Solid carbon and gaseous carbon dioxide are placed in a closed vessel and allowed to come to equilibrium at a high temperature according to this equation.



If the pressure on the system is increased and the temperature kept constant, what will be the result?

- (A) The amount of CO will increase and that of C and CO_2 will decrease.
- (B) The amount of CO will decrease and that of C and CO_2 will increase.
- (C) The amount of each of the three substances will increase.
- (D) The amount of each of the three substances will decrease.

19. Consider this reaction.



What is the equilibrium constant expression if the initial concentration of AB_3 is 0.1 M and the equilibrium concentration of A is represented by x ? Assume the initial concentrations of A and B are both zero.

- (A) $\frac{x-3x}{0.1-x}$ (B) $\frac{x-x^3}{(0.1-x)^3}$
 (C) $\frac{x \cdot x^3}{(0.1-3x)^3}$ (D) $\frac{x \cdot (3x)^3}{0.1-x}$

20. At 0 °C the ion product constant of water, K_w , is 1.2×10^{-15} . The pH of pure water at this temperature is

- (A) 6.88 (B) 7.00 (C) 7.46 (D) 7.56

21. What is the correct equation for the ion product constant of water at 25 °C?

- (A) $[\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$ (B) $[\text{H}_3\text{O}^+] + [\text{OH}^-] = 10^{-14}$
 (C) $\frac{[\text{H}_3\text{O}^+]}{[\text{OH}^-]} = 10^{-14}$ (D) $\frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]} = 10^{-14}$

22. What is the solubility product, K_{sp} , of $\text{Mg}(\text{OH})_2$ if its solubility in water is $1.6 \times 10^{-4} \text{ mol} \cdot \text{L}^{-1}$?

- (A) 1.6×10^{-11} (B) 2.6×10^{-8} (C) 3.2×10^{-4} (D) 4.1×10^{-12}

23. A saturated solution of MgF_2 contains $1.16 \times 10^{-3} \text{ mol}$ of MgF_2 per liter at a certain temperature. What is the K_{sp} of MgF_2 at this temperature?

- (A) 2.7×10^{-6} (B) 1.6×10^{-8} (C) 3.1×10^{-9} (D) 6.2×10^{-9}

24. What will be the result if 100 mL of 0.06 M $\text{Mg}(\text{NO}_3)_2$ is added to 50 mL of 0.06 M $\text{Na}_2\text{C}_2\text{O}_4$? Assume the reaction is taking place at 25 °C.

Substance	K_{sp} , 25 °C
$\text{MgC}_2\text{O}_4(\text{s})$	8.6×10^{-5}

- (A) No precipitate will form.
 (B) A precipitate will form and an excess of Mg^{2+} ions will remain in solution.
 (C) A precipitate will form and an excess of $\text{C}_2\text{O}_4^{2-}$ ions will remain in solution.
 (D) A precipitate will form but neither ion is present in excess.

25. The addition of solid Na_2SO_4 to an aqueous solution in equilibrium with solid BaSO_4 will cause

Substance	K_{sp} , 25 °C
$\text{BaSO}_4(\text{s})$	1.5×10^{-9}

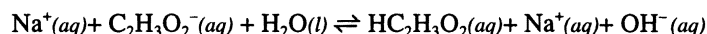
- (A) no change in $[\text{Ba}^{2+}]$ in solution.
 (B) more BaSO_4 to dissolve.
 (C) precipitation of more BaSO_4 .
 (D) an increase in the K_{sp} of BaSO_4 .

26. In a 0.050 M solution of a weak monoprotic acid, $[\text{H}^+] = 1.8 \times 10^{-3}$. What is its K_a ?

- (A) 3.6×10^{-2} (B) 9.0×10^{-5} (C) 6.7×10^{-5} (D) 1.6×10^{-7}

Equilibrium

27. A solution of sodium acetate in water is observed to become more alkaline as the temperature is raised. Which conclusion can be drawn? This is the equation for the reaction.



- (A) The forward reaction proceeds with evolution of heat.
(B) The forward reaction proceeds with absorption of heat.
(C) Acetic acid is less soluble in hot water than in cold water.
(D) At higher temperatures, $\text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{NaOH}(\text{aq})$ will occur.
28. A 0.15 M solution of a weak acid is found to be 1.3% ionized. What is its K_a ?
- (A) 1.3×10^{-2} (B) 2.0×10^{-3} (C) 1.1×10^{-3} (D) 2.6×10^{-5}
29. What is the value of the equilibrium constant K , for a reaction for which ΔG° is equal to -5.20 kJ at 50°C ?
- (A) 0.144 (B) 0.287 (C) 6.93 (D) 86.4
30. For the reaction



$\Delta H^\circ = +176$ kJ and $\Delta G^\circ = +91.2$ kJ at 298 K. What is the value of ΔG at 1000 K?

- (A) -109 kJ (B) -64 kJ (C) $+64$ kJ (D) $+109$ kJ

Answers to Study Questions

- | | | |
|------|------|-------|
| 1. D | 5. A | 9. A |
| 2. C | 6. C | 10. C |
| 3. B | 7. D | |
| 4. D | 8. B | |

Answers to Practice Questions

- | | | |
|-------|-------|-------|
| 1. D | 11. B | 21. A |
| 2. C | 12. B | 22. A |
| 3. A | 13. A | 23. D |
| 4. D | 14. B | 24. B |
| 5. D | 15. A | 25. C |
| 6. A | 16. B | 26. C |
| 7. D | 17. D | 27. B |
| 8. D | 18. B | 28. D |
| 9. C | 19. D | 29. C |
| 10. D | 20. C | 30. A |

Practice Questions

1. In which pair does the named element have the same oxidation number?

- (A) sulfur in $\text{H}_2\text{S}_2\text{O}_7$ and in H_2SO_4 (B) mercury in Hg^{2+} and in Hg_2^{2+}
 (C) oxygen in Na_2O_2 and in H_2O (D) cobalt in $\text{Co}(\text{NH}_3)_6^{3+}$ and in $\text{Co}(\text{NO}_3)_2$

2. In which compound does vanadium have the *lowest* oxidation state?

- (A) V_2O_5 (B) V_2O_3 (C) VO_2 (D) VO

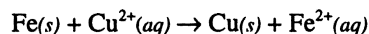
3. What is the oxidation number of chromium in $\text{Na}_2\text{Cr}_2\text{O}_7$?

- (A) +12 (B) +6 (C) +3 (D) -2

4. Which of these elements can display the largest number of different oxidation states?

- (A) aluminum (B) magnesium (C) manganese (D) mercury

5. Which statement is true for this reaction?

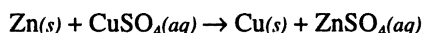


- (A) Cu^{2+} is oxidized. (B) Cu^{2+} gains in oxidation number.
 (C) Cu^{2+} is reduced. (D) Fe is reduced.

6. Which change requires an oxidizing agent to produce the indicated product?

- (A) $\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4\text{O}_6^{2-}$ (B) $\text{Zn}^{2+} \rightarrow \text{Zn}$
 (C) $\text{ClO}^- \rightarrow \text{Cl}^-$ (D) $\text{SO}_3 \rightarrow \text{SO}_4^{2-}$

7. Which statement is true for this reaction?



- (A) Metallic zinc is the reducing agent. (B) Metallic zinc is reduced.
 (C) Copper ion is oxidized. (D) Sulfate ion is the oxidizing agent.

8. In this reaction, which substance behaves as the oxidizing agent?



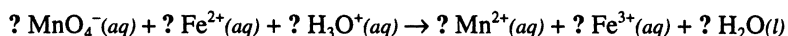
- (A) Pb (B) PbSO_4 (C) PbO_2 (D) H_2SO_4

9. Under certain conditions, H_2O_2 can act as an oxidizing agent; under other conditions, as a reducing agent. What is the best theoretical explanation for this?

- (A) H_2O_2 is a good bleaching agent.
 (B) Peroxides are stronger oxidizing agents than are oxides.
 (C) H_2O_2 will decolorize KMnO_4 solutions in the presence of an acid and will turn black lead sulfide to a white compound.
 (D) An atom within a compound can sometimes attain a more stable electronic structure either by gaining or by losing electrons.

Electrochemistry and Redox

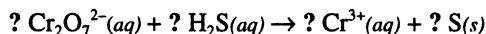
10. Balance this ionic equation for a redox reaction, using only whole number coefficients.



What is the coefficient for Fe^{2+} in the balanced equation?

- (A) 1 (B) 3 (C) 4 (D) 5

11. When this oxidation-reduction equation is balanced in acidic solution, using only whole number coefficients, what is the coefficient for $\text{S}(s)$?



- (A) 4 (B) 3 (C) 2 (D) 1

12. Which of these ions is the best reducing agent?

Standard Reduction Potentials	E°
$\text{Fe}^{3+}(aq) + e^- \rightarrow \text{Fe}^{2+}(aq)$	+0.77 V
$\text{Cu}^{2+}(aq) + e^- \rightarrow \text{Cu}^+(aq)$	+0.15 V

- (A) Fe^{3+} (B) Fe^{2+} (C) Cu^{2+} (D) Cu^+

13. A quantity of a powdered mixture of zinc and iron is added to a solution containing Fe^{2+} and Zn^{2+} ions, each at unit activity. What reaction will occur?

Standard Reduction Potentials	E°
$\text{Fe}^{3+}(aq) + e^- \rightarrow \text{Fe}^{2+}(aq)$	+0.77 V
$\text{Fe}^{2+}(aq) + 2e^- \rightarrow \text{Fe}(s)$	-0.44 V
$\text{Zn}^{2+}(aq) + 2e^- \rightarrow \text{Zn}(s)$	-0.76 V

- (A) Zinc ions will oxidize Fe to Fe^{2+} . (B) Fe^{2+} ions will be oxidized to Fe^{3+} ions.
(C) Zinc ions will be reduced to zinc metal. (D) Zinc metal will reduce Fe^{2+} ions.

14. Two metals are represented by the symbols L and Z. Which is the best oxidizing agent listed?

Standard Reduction Potentials	E°
$\text{L}^{2+}(aq) + 2e^- \rightarrow \text{L}(s)$	+0.76 V
$\text{Z}^+(aq) + e^- \rightarrow \text{Z}(s)$	+0.80 V

- (A) L (B) L^{2+} (C) Z (D) Z^+

15. Which combination of reactants will produce the greatest voltage based on these standard electrode potentials?

Standard Reduction Potentials	E°
$\text{Cu}^+(aq) + e^- \rightarrow \text{Cu}(s)$	+0.52 V
$\text{Sn}^{4+}(aq) + 2e^- \rightarrow \text{Sn}^{2+}(aq)$	+0.15 V
$\text{Cr}^{3+}(aq) + e^- \rightarrow \text{Cr}^{2+}(aq)$	-0.41 V

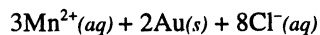
- (A) Cu^+ and Sn^{2+} (B) Cu^+ and Cr^{2+} (C) Cu and Sn^{2+} (D) Sn^{4+} and Cr^{2+}

16. What is the standard electrode potential for a voltaic cell constructed in the appropriate way from these two half-cells?

Standard Reduction Potentials	E°
$\text{Cr}^{3+}(aq) + 3e^- \rightarrow \text{Cr}(s)$	-0.74 V
$\text{Co}^{2+}(aq) + 2e^- \rightarrow \text{Co}(s)$	-0.28 V

- (A) -1.02 V (B) 0.46 V (C) 0.64 V (D) 1.02 V

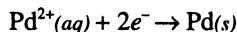
17. What is the standard cell potential, E° , for this reaction?



Standard Reduction Potentials	E°
$\text{Mn}^{2+}(aq) + 2e^- \rightarrow \text{Mn}(s)$	-1.18 V
$\text{AuCl}_4^-(aq) + 3e^- \rightarrow \text{Au}(s) + 4\text{Cl}^-(aq)$	+1.00 V

- (A) -0.18 V (B) -2.18 V (C) +2.18 V (D) +5.54 V

18. What is the standard cell potential, E° , for this half-reaction?



Standard Reaction Potentials	E°
$\text{Cu}(\text{s}) + \text{Pd}^{2+}(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + \text{Pd}(\text{s})$	+0.650 V
$\text{Cu}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Cu}(\text{s})$	+0.337 V

- (A) +0.987 V (B) -0.987 V (C) +0.313 V (D) -0.313 V
19. The standard cell potential, E° , for this reaction is 0.79 V.
- $$6\text{I}^-(\text{aq}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) \rightarrow 3\text{I}_2(\text{aq}) + 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$$
- What is the standard potential for $\text{I}_2(\text{aq})$ being reduced to $\text{I}^-(\text{aq})$ given that the standard reduction potential for $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ changing to $\text{Cr}^{3+}(\text{aq})$ is +1.33 V?
- (A) +0.54 V (B) -0.54 V (C) +0.18 V (D) -0.18 V
20. During the electrolysis of an aqueous solution of CuSO_4 using inert electrodes,
- (A) the anode loses mass and the cathode gains mass.
 (B) the mass of the anode decreases but the mass of the cathode remains constant.
 (C) the mass of the anode remains the same but the cathode gains mass.
 (D) the anode and the cathode neither gain nor lose mass.
21. Fluoride ion in aqueous solution are difficult to oxidize at the anode of an electrolytic cell because
- (A) the aqueous solutions of fluorides are nonconducting.
 (B) it is impossible to find the proper material from which to build the electrodes.
 (C) the fluorides are not very soluble.
 (D) oxygen is released from water in preference to fluorine.
22. Which products are formed during the electrolysis of a concentrated aqueous solution of sodium chloride?
- I. $\text{Cl}_2(\text{g})$ II. $\text{NaOH}(\text{aq})$ III. $\text{H}_2(\text{g})$
- (A) I only (B) I and II only
 (C) I and III only (D) I, II, and III
23. What mass of platinum could be plated on an electrode from the electrolysis of a $\text{Pt}(\text{NO}_3)_2$ solution with a current of 0.500 A for 55.0 s?
- (A) 27.8 mg (B) 45.5 mg (C) 53.6 mg (D) 91.0 mg
24. A current of 5.00 A is passed through an aqueous solution of chromium(III) nitrate for 30.0 min. How many grams of chromium metal will be deposited at the cathode?
- (A) 0.027 g (B) 1.62 g (C) 4.85 g (D) 6.33 g
25. A vanadium electrode is oxidized electrically. If the mass of the electrode decreases by 114 mg during the passage of 650 coulombs, what is the oxidation state of the vanadium product?
- (A) +1 (B) +2 (C) +3 (D) +4

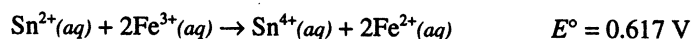
Electrochemistry and Redox

26. In which case would the *least* number of faradays of electricity be required for the liberation of 1.0 g of free metal?

Atomic Molar Masses	
K	39.1 g·mol ⁻¹
Na	23.0 g·mol ⁻¹
Cu	63.5 g·mol ⁻¹
Ag	107.9 g·mol ⁻¹

- (A) K from molten KOH
(B) Na from molten NaCl
(C) Cu from aqueous CuSO₄
(D) Ag from aqueous AgNO₃
27. How many minutes will be required to deposit 1.00 g of chromium metal from an aqueous CrO₄²⁻ solution using a current of 6.00 amperes?
- (A) 186 min (B) 30.9 min (C) 15.4 min (D) 5.15 min

28. Consider this reaction.



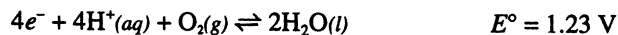
What is the value of E when $[\text{Sn}^{2+}]$ and $[\text{Fe}^{3+}]$ are equal to 0.50 M and $[\text{Sn}^{4+}]$ and $[\text{Fe}^{2+}]$ are equal to 0.10 M?

- (A) 0.069 V (B) 0.679 V (C) 0.658 V (D) 0.576 V

The Nernst Equation

$$E = E^\circ - \frac{0.0592}{n} \log(Q)$$

29. Consider this reaction.



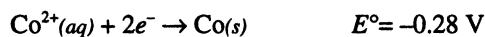
Which statement is true if the hydrogen ion concentration is initially at 1.0 M and the initial pressure of oxygen gas is 1.0 atmosphere?

- (A) Addition of a base should result in a value of E which is less than 1.23 V.
(B) $n = 1$, since one mole of oxygen is being considered.
(C) E is independent of the pH of the solution.
(D) $Q = \frac{[\text{H}_2\text{O}]^2}{[\text{O}_2][\text{H}^+]}$

The Nernst Equation

$$E = E^\circ - \frac{0.0592}{n} \log(Q)$$

30. Which of these changes will produce the most positive voltage for this half reaction in the direction written?



- (A) increasing the amount of solid Co
(B) decreasing the amount of solid Co
(C) increasing the concentration of Co²⁺
(D) decreasing the concentration of Co²⁺

Answers to Study Questions

1. A
2. C
3. C
4. B

5. B
6. B
7. C
8. C

9. B
10. B

Answers to Practice Questions

1. A
2. D
3. B
4. C
5. C
6. A
7. A
8. C
9. D
10. D

11. B
12. D
13. D
14. D
15. B
16. B
17. C
18. A
19. A
20. C

21. D
22. D
23. A
24. B
25. C
26. D
27. B
28. B
29. A
30. C

Practice Questions

- Which metal reacts with concentrated NaOH to produce hydrogen gas?
(A) aluminum (B) copper (C) iron (D) magnesium
- Which is the most satisfactory solvent for alumina, Al_2O_3 , in the electrolytic process for the preparation of aluminum?
(A) liquid water, H_2O (B) molten cryolite, Na_3AlF_6
(C) molten bauxite, Al_2O_3 with other oxides (D) molten sodium chloride, NaCl
- Which of these gases can be most efficiently collected by the displacement of water?
(A) ammonia (B) carbon dioxide (C) oxygen (D) hydrogen chloride
- Which gas is *least* soluble in water?
(A) H_2 (B) CO_2 (C) NH_3 (D) SO_2
- Which gas is the predominant nitrogen-containing pollutant in the atmosphere?
(A) N_2 (B) NH_3 (C) NO_2 (D) N_2O_5
- Helium is preferable to hydrogen for filling balloons because helium is
(A) less dense. (B) less expensive.
(C) easier to obtain. (D) chemically inert.
- Transition metals typically have all of these characteristics *except*
(A) forming colored compounds.
(B) showing a variety of oxidation states.
(C) having individual atoms with one or more unpaired electrons.
(D) having low melting points in the elemental state.
- Which compound is *not* expected to be colored?
(A) CuCl_2 (B) $\text{K}_2\text{Cr}_2\text{O}_7$ (C) TiO_2 (D) NiSO_4
- The members of which pair are most similar in color?
(A) chlorine, $\text{Cl}_2(g)$, and bromine, $\text{Br}_2(l)$
(B) silver chloride, $\text{AgCl}(s)$, and silver sulfide, $\text{Ag}_2\text{S}(s)$
(C) copper(II) chloride, $\text{CuCl}_2(s)$, and calcium chloride, $\text{CaCl}_2(s)$
(D) potassium permanganate, $\text{KMnO}_4(aq)$, and iodine, $\text{I}_2(g)$
- Which natural process tends to change the free nitrogen of the air into "fixed" nitrogen?
(A) lightning storms (B) photosynthesis
(C) respiration of animals (D) rusting of iron

11. When glucose, $C_6H_{12}O_6$, is completely oxidized with excess oxygen, what are the products?
- (A) H_2O_2 and CO_2 (B) H_2O and CO_2
(C) H_2O_2 and CO (D) H_2O and CO
12. Which ion may form a scummy precipitate with ordinary soap?
- (A) HCO_3^- (B) CO_3^{2-} (C) Na^+ (D) Ca^{2+}
13. Which statement best reflects the organization of the modern periodic table?
- (A) Elements are always arranged in order of increasing atomic weights.
(B) Metallic properties increase going from bottom to top in a family of elements.
(C) Nonmetallic properties tend to predominate for elements at the far right portion of the table.
(D) Each transition element is placed in the column of the main group element that it most closely resembles.
14. If gallium, atomic number 31, combines with selenium, atomic number 34, what is the most likely formula based on your knowledge of the periodic nature of the elements?
- (A) $GaSe$ (B) $GaSe_2$ (C) Ga_2Se (D) Ga_2Se_3
15. An element forms a basic oxide with the formula XO and a hydride with the formula XH_2 . The hydride reacts with water to give hydrogen. The element X could be
- (A) K (B) Ca (C) N (D) O
16. The elements X, Y, and Z form these compounds: XCl_4 , XZ_2 , and YO . What formula would you predict for the compound formed between Y and Z? Assume normal oxidation states for chlorine and oxygen.
- (A) YZ (B) YZ_2 (C) Y_2Z (D) YZ_3
17. The size of metal atoms
- (A) generally increases progressively from top to bottom in a group in the periodic table.
(B) generally increases progressively from left to right in a period.
(C) are smaller than those of the corresponding ions.
(D) do not change upon losing electrons.
18. Which of these elements has the *smallest* atomic radius?
- (A) fluorine (B) chlorine (C) bromine (D) iodine
19. Which of these atoms will be the *smallest*?
- (A) Si ($Z = 14$) (B) P ($Z = 15$) (C) Ge ($Z = 32$) (D) As ($Z = 33$)

20. The radii of the ions in this series decrease because

Ion	Ionic Radii
Na^+	0.095 nm
Mg^{2+}	0.065 nm
Al^{3+}	0.050 nm

- (A) the elements are in the same period.
 (B) the effective nuclear charge is increasing.
 (C) the atomic radius of Na decreases from Na to Al.
 (D) the first ionization energies increase from Na to Al.
21. The species F^- , Ne, and Na^+ all have the same number of electrons. Which is the predicted order when they are arranged in order of decreasing size (largest first)?
- (A) $\text{F}^- > \text{Ne} > \text{Na}^+$ (B) $\text{Ne} > \text{Na}^+ > \text{F}^-$
 (C) $\text{Na}^+ > \text{F}^- > \text{Ne}$ (D) $\text{F}^- > \text{Na}^+ > \text{Ne}$
22. Which isoelectronic ion is the *smallest*?
- (A) Al^{3+} (B) Na^+ (C) F^- (D) O^{2-}
23. What happens when a bromine atom becomes a bromide ion?
- (A) A positive ion is formed.
 (B) The bromine nucleus acquires a negative charge.
 (C) The bromide ion is larger than the bromine atom.
 (D) The atomic number of bromine is decreased by one.
24. Which ion has the largest radius?
- (A) Cl^- (B) F^- (C) K^+ (D) Cu^{2+}
25. Which of these elements has the highest electronegativity?
- (A) oxygen (B) iodine (C) cesium (D) lithium
26. Which element is more electronegative than arsenic and less electronegative than sulfur?
- (A) chlorine (B) phosphorus (C) tin (D) oxygen
27. List the elements Ca, Si, and K in order of *increasing* electronegativity.
- (A) $\text{Ca} < \text{K} < \text{Si}$ (B) $\text{K} < \text{Si} < \text{Ca}$
 (C) $\text{Si} < \text{Ca} < \text{K}$ (D) $\text{K} < \text{Ca} < \text{Si}$
28. Which pair of elements is listed in order of decreasing first ionization energy?
- (A) Na, Mg (B) Mg, Al (C) Al, Si (D) Si, P

29. When the species F^- , Na^+ , and Ne are arranged in order of increasing energy for the removal of an electron, what is the correct order?
- (A) $F^- < Na^+ < Ne$ (B) $F^- < Ne < Na^+$
(C) $Na^+ < Ne < F^-$ (D) $Ne < F^- < Na^+$
30. The first three ionization energies of an element X are 590, 1145, and $4912 \text{ kJ}\cdot\text{mol}^{-1}$. What is the most likely formula for a stable ion of X ?
- (A) X^+ (B) X^{2+} (C) X^{3+} (D) X^-

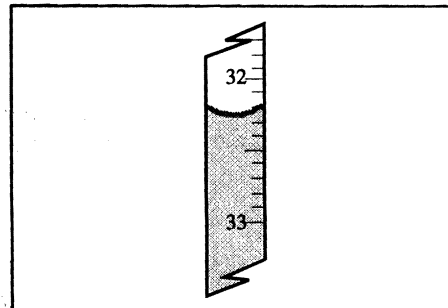
Answers to Study Questions

- | | | |
|------|------|-------|
| 1. D | 5. B | 9. D |
| 2. A | 6. A | 10. A |
| 3. B | 7. C | |
| 4. B | 8. A | |

Answers to Practice Questions

- | | | |
|-------|-------|-------|
| 1. A | 11. B | 21. A |
| 2. B | 12. D | 22. A |
| 3. C | 13. C | 23. C |
| 4. A | 14. D | 24. A |
| 5. C | 15. B | 25. A |
| 6. D | 16. A | 26. B |
| 7. D | 17. A | 27. D |
| 8. C | 18. A | 28. B |
| 9. D | 19. B | 29. B |
| 10. A | 20. B | 30. B |

5. What is the correct reading for the liquid in this buret?



- (A) 32 mL (B) 32.2 mL (C) 32.26 mL (D) 33.74 mL
6. To deliver a 25.00 mL liquid sample most precisely, which piece of glassware would you use?
- (A) volumetric flask (B) volumetric pipet
(C) graduated cylinder (D) eye dropper
7. A student determined the percent water in a sample. In four trials, values of 16.145%, 16.160%, 16.156%, and 17.279% were obtained for the percent water in a sample. What value should be used for the reported percent water?
- (A) 16.154% (B) 16.435% (C) 16.145% (D) 17.279%
8. Two students use the same standard mass of 1.0000 g to calibrate the same balance by taking three readings each.

Reading	Mike's Data	Jean's Data
1	1.0004 g	0.9996 g
2	0.9998 g	0.9994 g
3	0.9992 g	0.9995 g

Comparing the two sets of data,

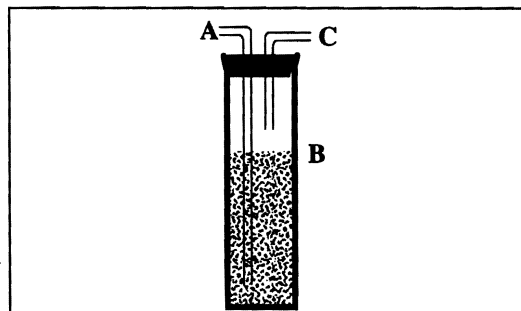
- (A) Mike's is more accurate, but Jean's is more precise.
(B) Jean's is more accurate, but Mike's is more precise.
(C) Mike's is both more accurate and more precise.
(D) Jean's is both more accurate and more precise.
9. Which procedure can be used to demonstrate experimentally that this reaction obeys the Law of Conservation of Matter?
- $$2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2\text{MgO}(s)$$
- (A) Take a mass of 1.0000 g of Mg ribbon, burn it in pure O_2 , and compare the mass of the product with the original mass of the Mg.
(B) Show that the sum of two atomic molar masses of Mg plus one molecular molar mass of O_2 is equal to two formula molar masses of MgO.
(C) Determine the mass of a sealed flash-bulb containing magnesium and oxygen, ignite the mixture, cool, and compare the final mass of bulb plus contents with the original mass of the bulb plus contents.
(D) Burn 1.0000 g of Mg ribbon in a tall beaker filled with air, scrape out all of the MgO formed, and compare with the original mass of the Mg.

Laboratory Chemistry

10. In which case could the given gas be identified by the reaction given?

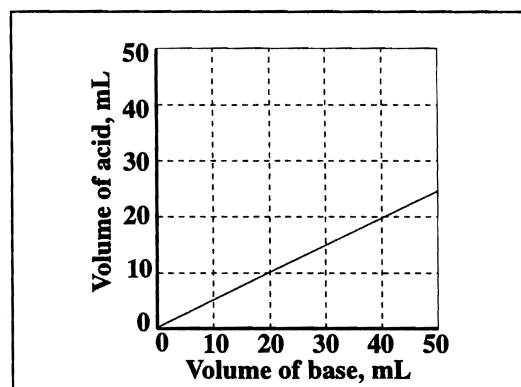
	Gas	Reaction
(A)	chlorine	will turn starch blue
(B)	ammonia	will turn moist red litmus paper blue
(C)	nitrogen	will support combustion
(D)	hydrogen chloride	will etch glass

11. Gas A enters the tube, passes through the contents of B, and emerges as gas C. If gas A is a mixture of carbon monoxide and carbon dioxide, and gas C is carbon monoxide, what is most likely substance in the tube?



- | | |
|---------------------------|-------------------------------|
| (A) concentrated HCl(aq) | (B) pure water |
| (C) phosphorous pentoxide | (D) sodium hydroxide solution |

12. A 1.0 M solution of an acid is used to titrate several different samples of a 1.0 M base solution. The graph shows the volume of acid needed to neutralize a given volume of base. Which reaction is consistent with the data?



- (A) $\text{HCl} + \text{KOH} \rightarrow \text{KCl} + \text{H}_2\text{O}$
- (B) $2\text{HCl} + \text{Ba}(\text{OH})_2 \rightarrow \text{BaCl}_2 + 2\text{H}_2\text{O}$
- (C) $\text{H}_2\text{SO}_4 + 2\text{KOH} \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$
- (D) $\text{H}_2\text{SO}_4 + \text{Ba}(\text{OH})_2 \rightarrow \text{BaSO}_4 + 2\text{H}_2\text{O}$

13. Which experimental evidence clearly illustrates the oxidizing property of chlorine, Cl_2 ?

- (A) its liberation at the anode during the electrolysis of sodium chloride solution
- (B) the reaction of chlorine with sodium bromide solution, resulting in the displacement of bromine
- (C) the use of chlorates as oxidizing agents in the manufacture of matches
- (D) its moderate solubility in aqueous solution

14. Consider this table of experimental data. In each case, two aqueous solutions were thoroughly mixed, and a small amount of organic solvent was then added, forming a separate layer in the reaction mixture.

Solution A	Solution B	Color of Organic Layer
NaI(aq)	Cl ₂ (aq)	purple
NaI(aq)	Br ₂ (aq)	purple
NaBr(aq)	I ₂ (aq)	purple

What conclusion can be drawn from these observations?

- (A) Iodine is able to displace bromide ion, but not chloride ion.
(B) Chlorine is able to displace both bromide and iodide ion.
(C) Both chlorine and bromine are able to displace iodide ion.
(D) Neither chlorine and bromine are able to displace iodide ion.
15. The best method of transferring a coarsely-powdered solid to a six-inch test tube is to
- (A) pour it through a thin-stemmed glass funnel.
(B) pour it from the lip of an evaporating dish.
(C) pour it from the bottle originally containing the solid.
(D) pour it from a creased square of paper.
16. A student combined 1 mL of 0.1 M solutions of nitrate salts of barium, calcium, magnesium, and strontium with several different testing reagents. The observations are recorded in this table.

Note: ppt = precipitate, rxn = reaction

Reagent	Ba(NO ₃) ₂	Ca(NO ₃) ₂	Mg(NO ₃) ₂	Sr(NO ₃) ₂
H ₂ SO ₄	white ppt	no rxn	no rxn	white ppt
Na ₂ CO ₃	white ppt	white ppt	white ppt	white ppt
(NH ₄) ₂ C ₂ O ₄	white ppt	white ppt	no rxn	white ppt
K ₂ CrO ₄	yellow ppt	no rxn	no rxn	no rxn

An unknown solution contains *two* of these four ions: barium, calcium, magnesium, and strontium. The unknown solution forms a white precipitate with H₂SO₄ and it does not react with K₂CrO₄. What conclusion about the two ions in the unknown can be drawn from this evidence?

- (A) It contains only strontium and barium ions.
(B) It contains only barium and calcium ions.
(C) It contains barium ion and *either* magnesium or calcium ions.
(D) It contains strontium ion and *either* magnesium or calcium ions.
17. You have a beaker nearly full of gasoline on your laboratory table and it catches on fire. What should you do?
- (A) Pour water into the beaker.
(B) Throw it into the sink.
(C) Cover the beaker with a wet cloth.
(D) Blow out the flames.

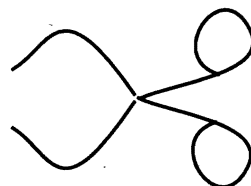
Laboratory Chemistry

18. Which of these pieces of apparatus would be best to use if you want to safely pick up a beaker containing hot liquids?

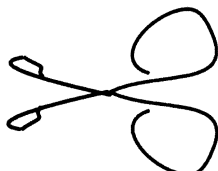
(A)



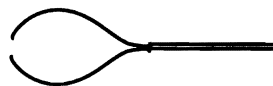
(B)



(C)



(D)



19. When diluting concentrated H_2SO_4 , the acid should be added to water because

- (A) concentrated sulfuric acid is a strong acid and can only be added to water.
- (B) concentrated sulfuric acid is such a good oxidizing agent that it is able to oxidize water.
- (C) concentrated sulfuric acid must be protected from air by a layer of water.
- (D) concentrated sulfuric acid and water generates considerable heat when in contact.

20. What should your response be if you enter an unfamiliar laboratory workspace and see this sign?



- (A) You should immediately evacuate the premises and then call for OSHA.
- (B) You should put on full protective clothing, including goggles, respirator, and body suit.
- (C) You should be aware that chemical waste is being processed in the immediate area.
- (D) You should pay careful attention to your surroundings for potential hazards.

Answers to Study Questions

- 1. A
- 2. D
- 3. D
- 4. D

- 5. C
- 6. A
- 7. B
- 8. B

- 9. A
- 10. D

Answers to Practice Questions

- 1. A
- 2. D
- 3. B
- 4. B
- 5. C

- 6. B
- 7. A
- 8. A
- 9. C
- 10. B

- 11. D
- 12. C
- 13. B
- 14. C
- 15. D

- 16. D
- 17. C
- 18. B
- 19. D
- 20. D